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CHEMISTRY PRACTICALLY TAUGHT

NEW EDITION OF WATTS'S DICTIONARY OF CHEMISTRY.

WATTS'S DICTIONARY OF CHEMISTRY.

REVISED AND ENTIRELY RE-WRITTEN

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To be published in 4 vols. 8vo. Vol. I. royal 8vo. price 42*s*.

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The great advances in Chemistry made within recent years have rendered it almost imperative that a new edition of Watts's Dictionary should be prepared. Not only has an immense number of descriptive facts been added to Chemistry in the quarter of a century which has elapsed since the publication of the first edition began, but the whole aspect of the science, and its relations to other branches of natural knowledge, have greatly changed in that time. The changes have been so great and far-reaching that a new edition of a Chemical Dictionary which saw the light twenty-five years ago must needs be a new book.

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The Editors have obtained the help of many chemists—English, American, and foreign—of acknowledged repute.

London: LONGMANS, GREEN, & CO.

THE

FUNDAMENTAL PRINCIPLES
OF
CHEMISTRY

PRACTICALLY TAUGHT, BY A NEW METHOD

BY

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BE TAUGHT' 'A TREATISE ON FUEL, SCIENTIFIC
AND PRACTICAL' ETC.

LONDON
LONGMANS, GREEN, AND CO.
AND NEW YORK : 15 EAST 16th STREET
1888

M. F. W.

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PREFACE.

THE very favourable reviews of my work on 'Education, Scientific and Technical,' especially the portion relating to the teaching of Chemistry, were in a measure the cause of my writing the present work. I have long held that chemical works intended for beginners are unsuitable as educational works; if these books extend only to a few pages the arrangement and construction is the same as that adopted in Gmelin's great work of *reference* on the science, which extends to eighteen large volumes: the plan is encyclopædic—excellent for a book of reference, unsuitable for an educational work. In this system the facts are unclassified; the laws, the highest generalisations, are placed apart from the facts; and no plan for teaching the language of the science, which requires to be taught like any other language, is given beyond a few general observations.

These views I put forward more than twenty years ago in the preface to another work, and I am glad to find that, in process of time, other chemists are beginning to have similar views on this class of books. Professor Josiah P. Cooke, of the Harvard University, thus comments upon them, in the introduction to the new edition of his work, 'The New Chemistry:' 'To the great mass of learners the study of these text-books is uninteresting and profitless; for before the student is made familiar, through long laboratory practice, with the materials and processes described, such a book is little more to him than a catalogue of names, to which he attaches no significance. . . . On this plan' (the plan of his 'New Chemistry') 'the student's interest can be sustained to the end, which is hardly possible in following through the weary catalogue of elementary substances given in the usual text-books, involving a repetition of details as profitless to the general student as it is tedious and uninteresting.'

The plan of this book is, I am confident, more in harmony with the laws of thought than the chemical text-books that are in use in schools, and for teaching junior students generally. Many of these books are little more than mere compilations of unclassified facts; hence the reason is not cultivated, and consequently the science is not attained; for unless the first is reached, the second is not gained.

As the student cannot be properly instructed in chemistry without having previously acquired a knowledge of the physical properties of matter and the physical forces, I have commenced with a course of physics suitable for the course of pure chemistry given in the after part of the work. In the teaching of this portion of the work the exercises, illustrations, &c., have been selected to bear on the after course, and on chemical operations generally; thus, in explaining porosity, filtration is illustrated and taught practically; the collecting and storing of gases under impenetrability; the determination of boiling-points, fractional distillation, &c., under heat; the action of charcoal and dyeing under adhesion, &c.; so that when the purely chemical portion of the work is reached the student will not be perplexed and impeded when, in describing chemical operations and chemical properties, reference has to be made to physical properties and physical forces, as he has already become practically acquainted with them.

The principles, both in the course on chemical physics and on pure chemistry, are taught by experimental and arithmetical exercises and examination questions; the questions, both on physics and chemistry, which I have not framed myself, have been taken mainly from the examination papers of the London University, the Royal University of Ireland, and Owens College, Manchester; and they and the exercises have been selected so as to impart a wide and sound knowledge of the different subjects dealt with: answers to many of the exercises are given at the end of the work. The language of the science and the mode of expressing chemical changes are taught in a series of progressive exercises, which will give the student a complete mastery over it. It being all-essential that the student should at an early period of his studies become acquainted with the language, this is essentially necessary, for without a knowledge of it he would be unable to think and reason about the science, because language is the atmosphere in which thought lives. It is for this reason the student is directed to pass over portions of Chapters VI., IX., XI., XII., XIII., until he has reached and studied Chapter XIV., as he has then completed the study of the language, and he is thus enabled to study more intelligently the more advanced portions of these chapters.

This book is complete on all the subjects it treats upon ; but should its sale justify it, a second volume will be published giving the advanced portions of the science which have not been dealt with in this.

I am greatly indebted to two friends and former students, Mr. James Roberts, LL.B., Barrister-at-Law, and Mr. Thomas Bayley, consulting chemist, in kindly assisting me in eliminating errors.

ROBERT GALLOWAY.

LONDON : *February* 1888

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Page 145, line 29, *for* $(P-p)=\frac{760}{4-1}$ *read* $(P-p)\frac{760}{H-h}$.
,, 216, ,, 30, *for* 644 *read* 643.
,, ,, ,, 31, *for* 670 *read* 673.
,, ,, ,, 36, *for* 721 *read* 724.
,, 226, ,, 26, *for* 700 *read* 703.
,, ,, ,, 27, *for* 661 *read* 664.

THE FUNDAMENTAL PRINCIPLES

OF

CHEMISTRY.

CHAPTER I.

MATTER—STATE OF AGGREGATION OF BODIES—PHYSICAL PROPERTIES OF BODIES—EXPERIMENTS—MOLECULAR FORCES—COLLECTING AND STORING GASES—FITTING UP CHEMICAL APPARATUS—QUESTIONS.

1. THOSE who take up this book with the intention of studying it, will, no doubt, have heard the word **matter** employed in conversation, and they will also have without doubt met with it in some of the books they have read; and most probably they themselves will have used it both in conversation and in writing; they must therefore have attached some meaning to this important word, for it is a very *important word*, especially for what it signifies in **Natural Science**. I wish those studying this book to reflect for a few moments before going further as to the meaning they attached to it when they or others employed it. Like many other words it is employed both in ordinary and scientific language. The following are a few of its meanings in ordinary language, and they are given in order to impress on the young student, that whilst a word has many different significations in ordinary language, it has one, and only one, in the particular department of science in which it is employed as a scientific term; hence one word cannot be substituted for another in scientific, as it frequently can in ordinary, language to convey the same idea; just as, for instance, if we speak of any tool, say a spade, hammer, &c., we cannot substitute any other word for the word spade or hammer,

as no other word would convey to other minds the particular tool we meant. The student will, we hope, gather from these marks the imperative necessity of learning, in order to become acquainted with a science, the *exact* and *precise* meaning of every scientific word and phrase employed.

2. The following are a few of the meanings of the word 'matter' as given in dictionaries :—

1. The contents of any speech or writing; the thing said; the meaning; sense. (2) Good sense; substance, as opposed to empty verbosity or frivolous jesting. 'To speak all mirth and no matter.' (3) Subject; thing treated; that about which we write or speak. (4) Affair; business; event; course of things; as 'Matters have succeeded well thus far'; 'Observe how matters stand'; 'Thus the matter rests at present'; 'Thus the matter ended.' (5) In logic that which forms the subject of any mental operation, as distinguished from the form. (6) Body; substance extended; anything perceptible by any of the senses; that of which the whole sensible universe is composed.

3. The last meaning is the only one in which it is used in Natural Science, or, in other words, when its meaning is confined to the **Material World**. For when employed in Natural Science it includes, but is restricted to, everything which possesses **weight** and occupies **space**; these are termed the essential properties of matter, since they suffice to define it.

4. A young child knows from observation that *all* matter is not of the same kind, but that there are many different kinds; there are, for example, gold, silver, iron, copper, sulphur, common salt, chalk, and a very large number of other kinds of matter. Now each of these different kinds of matter is termed a *substance* or *body*; and each substance exists in one or other of the three states of *aggregation* in which matter can exist, viz. the *solid*, the *liquid*, and the *gaseous* state; some substances exist naturally in only one of these states, as, for example, gold, silver, or iron in the solid state; but most of them, and perhaps all of them, can be made to pass into one or both of the other states in which they do not naturally exist, by the application of heat, cold (diminution of heat), or pressure; or by a combination of pressure and one of the other two agents, as we shall learn further on. Water, for example, exists naturally in all the three states; in all climates it exists in the liquid state; and in all climates it exists in the gaseous or vaporous state in the air; and in the Arctic regions, and in temperate climates

in winter, it exists in the solid state as ice and snow. Some of the water in the gaseous state in the air passes, or becomes converted, into one of the other two states of aggregation, owing to a *diminution* of temperature in the atmosphere; it then appears in the form of dew, or comes down to the earth in the form of rain—the liquid state, or of hail or snow—the solid state.

5. It has already been stated that *all* substances, whether existing in the solid, the liquid, or the gaseous state, occupy a certain space; the amount of space a substance occupies is termed its **volume.**

6. Every particle of matter attracts every other particle of matter, whether the matter exists in the solid, the liquid, or the gaseous state, not only in the world we dwell in, but throughout the entire universe; this attraction is called the **Law of Gravitation.** This force or attraction operates at the greatest conceivable distances, and is common, as we have stated, to the three forms of matter, and as each particle attracts, the *mass* containing the greatest number of particles must consequently exert the greatest attractive force; hence it is that all bodies, with the exception of the celestial ones, when at some distance from the earth's surface and *unsupported*, fall or descend to it. One of the effects of this attraction is the downward pressure which substances exert on whatever supports or prevents them from falling to the earth; *this pressure is termed their weight.*

7. The word 'particle' which has been employed has several different meanings in ordinary language, but it is only necessary to give the one which corresponds with its scientific meaning, viz. 'a minute part or portion of matter, the aggregation of which parts constitutes the whole mass;' this corresponds with its scientific meaning, for the smallest portion of matter which is *mechanically divisible* is termed a particle. But it is believed by scientific men that there are more minute portions of matter than particles; these minute portions are *indivisible*; the largest of these *mechanically* indivisible portions are termed *molecules*; the molecules are *chemically divisible*; the still smaller portions, which are both *mechanically* and *chemically indivisible*, are termed *atoms.*

8. Before going further we must inform the student that the terms particle and molecule have been used indiscriminately, and in some scientific works at the present time the word particle is sometimes used in place of the word molecule ; but in this work the two words have separate and distinct meanings; particle being employed to signify the smallest portion of matter that is *mechanically divisible* ; molecule the smallest portion which can exist in the *free state*, and which is *mechanically indivisible* but *chemically divisible*.

9. It will be evident to the student from what has been stated that the minutest particle is an aggregation of molecules, as a mass of matter is an aggregation of particles. Science cannot at present inform us how many molecules are required to form the minutest particle ; but, to form an idea of the size of molecules, Sir William Thompson gives this illustration. Imagine a drop of rain, or a glass sphere the size of a pea, magnified to the size of the earth, the molecules in it being increased in the same proportion ; the magnified structure would be coarser-grained than a heap of small shot, but probably less coarse-grained than a heap of cricket-balls. The number of molecules of gas in a cubic centimetre (0·06103 cubic inch) of air is calculated at twenty-one trillions. Molecules and atoms are treated of in a subsequent chapter.

10. We will here cite a few examples to show to what an almost incredible extent matter admits of being subdivided ; and the minuteness of some organised bodies.

1. 'Wollaston succeeded in obtaining threads of platinum of a diameter not exceeding $\frac{1}{3000000}$ of an inch. The method which he employed for preparing them consisted in drawing a silver wire with a platinum core, and dissolving the shell of silver in nitric acid. In this way threads can be obtained so fine that they are actually invisible to direct view ; their existence can only be detected by the aid of certain special optical phenomena. In the art of beating gold, leaves are obtained whose thickness cannot exceed $\frac{1}{200000}$ of an inch. A square inch of this leaf would weigh less than the $\frac{1}{20000}$ of an ounce, and as a square whose side is $\frac{1}{250}$ of an inch is visible to the naked eye, it follows that this square inch of leaf contains more than 60,000 visible parts.'

2. 'According to Biot, the thread by which a spider lets itself down is composed of more than 6,000 single threads. But more surprising than all is the microcosm of organised nature in the infusoria. Of these creatures, which for the most part we can see only by the aid of the microscope, there exist many species so small that millions piled on each other would not equal a single grain of sand, and thousands might swim at once through the

eye of the finest needle. [illegible] exist in [illegible] Ehrenberg's investigations [illegible] the smooth [illegible] By microscopic measurement [illegible] 41,000 millions [illegible] weighs [illegible] grain [illegible] which would [illegible] Conceive further [illegible] tions have proved [illegible] our forming the most remote conception of the [illegible] forms.

3. "The two following [illegible] perception [illegible] the presence of matter [illegible] could afford us the slightest [illegible]

4. "It is well [illegible] supply the air [illegible] cient number of particles [illegible] near mass may be [illegible] smell. The [illegible] must volatilise from the mass [illegible] perceptible in all directions [illegible] their *minimum*. [illegible] rated in a [illegible] length of time."

11. These and numerous other examples, which might be given, would, if there were no [illegible] that matter can be infinitely divided; [illegible] tation of chemical phenomena [illegible] a *limit*, as we have previously [illegible] plain further on in the *structure* of matter.

12. When particles cohere together then, as has been stated, a mass of matter is formed, and the more numerous the particles cohering, the larger, of course, the volume of the mass; a piece of lump sugar, for example, is an aggregation of a number of particles of that substance. In some substances the particles are, as it were, indiscriminately [illegible]; in other substances they are arranged in a crystalline form, as seen in sugar, alum, epsom salts, &c.

13. The particles cohere together in different solid substances with very different degrees of attractive force; some solids, for instance, are very easily pulverised, that is their particles are very easily separated; other solids, on the contrary, resist pulverisation by ordinary means. The particles of the substances employed as building materials, as stone, iron, &c.,

are united very firmly together; hence the great strength of these substances; and this is the case with all solids whose particles are, like the solids named, firmly bound together. We may here state, what has already been implied, that *each particle* of a pulverised body possesses the *same properties or qualities* as the entire mass did before pulverisation.

14. In some cases if the parts of the same substance which have been divided are brought into *very intimate* contact they will unite; thus on pressing together two clean, smooth, and freshly cut surfaces of lead, they cohere, and a force of some pounds is required to separate them. In the same way, two perfectly polished plates of glass cohere, sometimes so completely that they may be cut and worked as a single piece. This has not unfrequently happened in plate-glass manufactories. No doubt two pieces of all substances might be made to cohere, as lead does, if they could be brought into sufficiently intimate contact. When it is desired to unite two pieces of a metal like iron, which cannot be made to unite by simply pressing them together, the cohesion between the particles of each piece is first weakened by heating them; the consequence is they become quite soft, and can be readily bent; when they are pressed or hammered together in this state they cohere and form one piece; this operation is termed *welding*.

15. However compact a substance in the solid state may appear, its molecules are not in actual contact, there are *spaces* or *pores*, as the spaces are termed, between the molecules. The Florentine Academicians in 1661 proved the existence of pores between the molecules of so compact a solid as gold; they proved this by filling a thin globe of gold with water, and after hermetically sealing its orifice subjected the globe to pressure; the result was, that the water was forced through the pores of the gold, and appeared in the form of dew, on the outer surface of the globe. In the building of the Britannia bridge it was also found that the water used in the enormous hydrostatic presses employed to raise the tube, was freely forced through the pores of the thick iron cylinders of the press.[1]

---

[1] 'No decisive proof of the porosity of vitreous bodies, such as glass, seems yet to have been obtained. That they form almost a solitary class of excep-

16. There are two kinds of pores, viz., *sensible* pores and *physical* pores; the sensible pores are apparent to the eye, either alone or aided by the microscope. Filtration is an application or utilisation of the porosity of certain substances, as paper, woven fabrics, stones, &c.; the sensible pores of these substances are sufficiently large to allow liquids to pass through them, but small enough to arrest the passage of *solid* particles which may be suspended in the liquid.

17. The physical pores are not apparent to the sight, and are the spaces or intervals between the molecules.

18. The space between the molecules can be diminished to a greater or less extent, in other words the molecules can be forced nearer to each other by the action of external pressure. This property is called *compressibility*, and is both a consequence and a proof of their porosity.

Experiments on Porosity.

1. Immerse a piece of chalk in water, air-bubbles will at once begin to ascend to the surface of the liquid, in consequence of the water displacing the air in the pores of the chalk.

2. Pour some water on a piece of marble; the water will rest on the surface, it will not enter into the pores, unless great pressure is used and for a sufficient length of time. Yet chalk and marble are formed of the same constituents; they differ only in their physical properties; and this difference is chiefly, if not entirely, due to the cohesiveness being less between the particles of the chalk than between the particles of the marble, and therefore the greater porosity of the former than of the latter substance.

19. Different liquids possess different degrees of facility for penetrating the pores of solid bodies; for instance, if polished marble is wetted with water very little, if any, enters its pores, while oil, on the contrary, is rapidly absorbed by it; and a bag of cambric or gauze, whose pores are visible to the eye, holds mercury securely, until sufficient pressure is applied to force out the liquid.

20. As filtration is in constant use in chemical laboratories, and as it is also a utilisation of the porosity of substances, we will now describe the operation, and illustrate it with experiments. Filtration is employed for separating liquids from

tions to an otherwise general rule seems highly improbable. It is most likely that vitreous bodies have not yet been proved to be porous *only* because we have not yet discovered the proper mode of testing them for this property.

solids, and usually is accomplished by passing the mixtu solid and liquid matter through filtering paper (a paper for the purpose) of proper size and shape, supported funnel. The pores of the paper permit the fluid, and the *dissolved* in it, to pass, whilst the undissolved solid matter b unable to pass through the pores, remains behind on the fi Sometimes, however, the undissolved solid matter is in su finely divided state that the filter does not at first intercep the solid particles, some portion passing through the pore the filter ; in such cases the imperfectly filtered portion be returned to the filter until it becomes *perfectly* bright clear.

21. If any *coloured* solid dissolves in the liquid the liq even when perfectly filtered, will not be colourless, but *colou*

22. To prepare a filter, cut the filtering paper in circ form, and fold it so as to form a half-circle, and then a fold it at right angles to the first fold (fig. 1); the paper ought to fall wholly within the funnel (fig. 2). The paper, when placed within the funnel, must be moistened with water, and then the liquid to be filtered must be gradually poured on. The funnel is supported in one of the rings of the retort-stand (fig. 2).

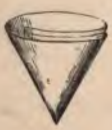

FIG. 1.

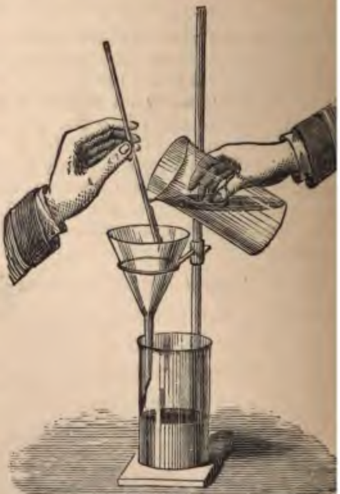

FIG. 2.

23. The liquid which passes through the filter is called *filtrate.* When the solid particles are very heavy, the supern tant liquid can be separated from the solid, without passing through a filter. This is accomplished by simply inclining t vessel, so as to allow the fluid to pass away unattended by the 'undissolved solid'; this is called decantation; or the fluid may be removed by a syphon.

Experiments on Filtration.

3. Mix some powdered chalk with water; then pass the mixture through a filter; if the filtrate comes through turbid at first (not clear), return it to the filter as often as is necessary to obtain it perfectly clear and bright.

4. Mix some soil with water; a portion of the soil will dissolve in the water, and a portion will remain undissolved; subsequently filter the solution; the filtrate will most probably be *coloured*, on account of some of the coloured components of the soil having dissolved in the water.

24. We have yet to mention a remarkable class of illustrations proving the porosity of solid bodies. Not only can gases pass through unglazed pottery, but hydrogen, the lightest of known gases, can pass through platinum and iron when they are red-hot; but we must reserve any further remarks on the passage of gases through metallic and non-metallic septa until we treat on the motion of gases in Part II.

25. The cohesive attraction between the molecules of substances varies not only in different substances, that is in substances of unlike constituents, but also, as shown by Expt. 2, in some substances composed of the same constituents; and an alteration in the molecular attraction can in some solids be produced by artificial means, the difference in the mechanical properties of tempered and untempered steel, for example, is owing to the difference in the molecular attractions produced by the tempering.

26. It is owing to a natural difference in the molecular attraction in some solids, and to the alteration which can be effected in the molecular arrangement in others by the action of physical agents, that solids vary, or can be made to vary, in the properties of *hardness*, *tenacity*, *ductility*, and *malleability*. Silver, gold, and platinum, which are the most ductile and malleable of the metals, can be drawn out into the form of wire, or beaten out in the form of thin leaves to an almost inconceivable extent without any rupture or breach of continuity between the molecules taking place. Platinum, which is the most ductile of metals, has been drawn out into wire so fine that it was only 0·00003 of an inch in diameter; and gold, which is the most malleable of metals, has been beaten into leaves so fine that they were only about $\frac{1}{300,000}$th of an inch thick.

27. Ductility and malleability are not the same property or quality; for the same metals are not always both ductile and malleable, or do not at least possess these properties in an equal degree. Iron possesses ductility in a much greater degree than it possesses malleability; for it admits of being drawn into extremely fine wire, though it cannot be beaten into extremely thin plates. Tin and lead, on the other hand, are highly malleable; they can be beaten into extremely thin leaves; but they are not ductile, since they cannot be drawn into fine wire.

28. A property the very opposite of malleability and ductility is *brittleness*; the molecules of brittle substances do not, as in malleable and ductile solids, admit of any very considerable displacement without losing their cohesiveness; in other words a brittle substance cannot be much altered in form without breaking.

29. We must now explain the meaning of the word **properties** we have been employing. By the term properties, as applied to bodies, we understand the different ways in which bodies present themselves to our senses, and we distinguish *general* from *specific* properties. The former are shared by all bodies, and among the most important are *impenetrability*, *extension*, *divisibility*, *porosity*, *compressibility*, *elasticity*, *mobility*, and *inertia*.

30. Specific properties are such as are observed in certain bodies only, or in certain states of these bodies; such are *solidity*, *fluidity*, *tenacity*, *ductility*, *malleability*, *hardness*, *transparency*, *colour*, *&c.*

31. We have so far, in speaking of the cohesion of the molecules of bodies, confined and limited our descriptions and illustrations to solid bodies; we have now to notice how far this attraction between the molecules of the *same substances* operates in bodies that are liquid, and in bodies that are gaseous at ordinary temperatures.

32. As regards the particles of liquids, they have, as compared with those of solids, no fixed position, on account of the less cohesive attraction between them; hence they glide or pass over each other with the greatest ease, being displaced by

slightest force ; liquid bodies on this account are capable
ssuming the shape of any vessel in which they are placed.
there is nevertheless a greater or less cohesion between the
ecules of liquids, which causes them to be more or less
ous. But the cohesive power between the molecules of
ids is more strikingly shown in the rounded form assumed
detached drops ; this is very beautifully exemplified in rain-
ps as they fall through the air, and in drops of dew ; in these
ll masses, water, as is seen, assumes the spheroidal shape.

33. The difference in the form assumed by large and small
ses of liquids is thus explained—in large masses the force
ravity overcomes the cohesive force ; hence liquids acted
n by the former force have no special shape ; they take
of the vessel in which they are contained ; and the free
ace of the liquid is a horizontal plane perpendicular to the
e of gravity. But in smaller masses cohesion overcomes
force of gravity, and therefore in minute masses liquids
me the spherical or globular form.

periments Illustrating the Tendency of Liquids in Small Masses to Assume the Spherical Form.

. Introduce by means of a pipette (fig. 3) a drop of oil
the middle of some alcohol which has been diluted with
r to precisely the same specific gravity as the oil ; the
liquids will not mix with or dissolve each other, but
il will remain suspended in the diluted alcohol in the
of a perfectly spherical mass.

. Pour into a narrow-necked bottle a saturated solution of
sulphate, add a few drops of carbon disulphide, coloured
me iodine dissolved, which will float on the surface of
zinc solution. Then carefully add some distilled water
to make it rest on the surface of the zinc solution and
nix with it ; the carbon disulphide will, if these conditions
been obtained, collect in the form of a flattened spheroid,
nting the appearance of blown coloured glass, and will
rger than the neck of the bottle if a sufficient quantity of
arbon compound has been employed.

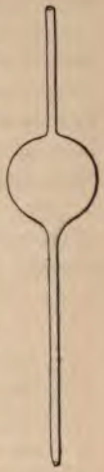
FIG. 3.

34. The Florentine Academicians made the ex-
ment which has already been described (par. 15)
for ascertaining whether solids were or were not
us ; but whether liquids were or were not com-
sible ; the liquid they operated upon, as already
d, was water. As the globe was completely filled with

water, any pressure on it they expected would, by altering its shape and diminishing its capacity, either compress the liquid if it were compressible, or, if it were incompressible, would burst the globe. Neither of these expected results occurred; but, contrary to the expectations of the experimenters, the experiment proved that gold was porous.

35. The Academicians inferred, from the results of the experiment, that water was not compressible, and this opinion was accepted as correct for some time afterwards; but later researches have proved that all liquids are compressible, but only to a slight degree as compared with either solids or gases. Thus, for a pressure equal to that of the atmosphere—15 lbs. per square inch—mercury experiences a compression of 0·000005 part of its original volume, water a compression of 0·00005, and ether a compression of 0·000133. The compressibility of liquids is, however, greater the higher their temperature; thus that of ether at 14 °C. is one-fourth greater than its compressibility at 0° C. The apparatus employed for measuring the compressibility of liquid is named the **piezometer**, the term being derived from two Greek words, one signifying 'I compress,' and the other 'measure.'

36. As liquids are compressible their molecules, as in the case of solids, are not in actual contact, but are separated by spaces termed pores. The pores in liquids are not so obvious as they are in many solids, but it is none the less certain that they exist in liquids as in solids; this is proved not only from their being compressible; but also from the fact that two liquids that can mix together will, when added together, often occupy a less volume than the two occupied before being mixed.

Experiment on the Porosity of Liquids.

7. Half-fill with water a glass bottle, in shape like the specific gravity bottle (fig. 47), and fitted with a ground stopper, then pour some alcohol carefully down the inside of the vessel, so as not to mix the two liquids. The bottle should be so filled that a few drops of alcohol are pushed out when the stopper is inserted; this must be done in such a careful manner that no air-bubbles enter the bottle. Raise the bottle between the thumb and fingers, press the forefinger firmly on the stopper, and then invert the bottle several times so as to mix the two liquids together. When they are thoroughly mixed, it will be seen, on placing the bottle on the table, that

he volume of the mixed liquids is less than they occupied before they were mixed, hence proving their porosity. The same phenomenon may be oberved in mixing many other liquids. If the water is coloured red with solution of aniline-red (magenta), it will be seen whether the alcohol has een successfully introduced into the bottle.

37. When the external compressing force is removed, liquids always regain exactly their original volume. Bodies which recover their original form *perfectly* on the removal of the external pressure are said to be perfectly elastic under compression; if they only partially recover their original shape, their elasticity is said to be imperfect, and they remain distorted permanently. *All* liquids and gases are *perfectly elastic*, but no solid is perfectly elastic, this property being always confined in solids within *certain limits*. Imperfect elasticity gives rise to the remarkable properties of *ductility* and *malleability* in some solids, properties we have already noticed.

38. The elasticity we have been speaking about is called the **elasticity of compression.** Two other kinds of elasticity are recognised, viz., the **elasticity of extension,** and the **elasticity of torsion.** Liquids and gases possess only the elasticity of compression; the other two kinds are possessed by solid bodies only. The student is referred for further information on these two latter kinds of elasticity to works on natural philosophy.

39. There is no cohesive attraction between the particles of gaseous bodies; they actually repel each other, and to such an extent does this repulsion extend, that the particles of a gas not confined in a close vessel, or *subjected* to some external pressure, would disperse in all directions, owing to their mutual repulsion; they, therefore, exert a pressure in an outward direction upon the sides of the vessels which contain them; and the pressure is greater or less according as the elasticity of the gas is increased or diminished.

40. Owing to the absence of any cohesive attraction between their particles, gases are much more compressible than either solids or liquids; by sufficient pressure they may be made to occupy ten, twenty, or even some hundred, times less space than they occupy under ordinary circumstances. There is, however, *a limit* beyond which, when the pressure is increased, they are converted into the *liquid* state.

45. Reacting against this and the other molecular forces—these we shall notice further on—is the repulsive power of heat; and upon the relative intensity of these two opposing forces—cohesion and repulsion—depends the solid, fluid, or gaseous condition of matter. When the cohesive force prevails, bodies exist in the solid state; when the repulsive power is in the ascendant, matter, as we have learned, assumes the gaseous form; and when the two forces are balanced, the liquid state is the result.

46. We must here explain the meaning of the term 'force': 'When a body is at rest it has no *innate* power of changing its state of rest; when it is in motion it has no innate power of changing its state of uniform motion in a straight line. This property of matter is termed its *inertia*. Any cause which sets a body in motion, or which changes the magnitude or direction of its velocity if in motion, is a *force*. *Gravity*, *friction*, *elasticity* of springs or gases, electrical or magnetic, *attraction* or *repulsion*, etc., are *forces*. All changes observed in the motion of bodies can be referred to the action of one or more forces.'

47. Before treating at greater length on the repulsive power of heat, it is desirable to notice another general property of matter, viz. impenetrability. Impenetrability is that property of matter which prevents two substances from occupying the same space at the same time; we cannot, for instance, cause one particle or mass of matter to occupy the place of another particle or mass without *moving* or *displacing* it, and the opposition met with in displacing it is termed its *resistance*. The most cursory observation is sufficient to inform anyone that solid and liquid bodies possess this property; but it is not so readily apparent that gaseous bodies possess this property; because matter in the gaseous state can be displaced without sensible effort, and the gases most commonly met with are colourless, and therefore not evident to the sight.

48. The impenetrability of the air is easily shown by inserting a tumbler, mouth downwards, into a vessel containing water; the water will not enter the tumbler, except the small quantity due to the compression of the air, on account of the tumbler being already filled with air. The diving-bell is an exceedingly interesting application of this property; the bell may be said to be a large tumbler, which is inserted mouth downwards in water: the diver is perfectly safe in the bell owing to the impenetrability of the air: the only disagreeable feeling he experiences, as the bell descends in the water, is due to the compression of the air.

Experiments Illustrating the Impenetrability of Gaseous Bodies.

11. Bend a glass tube in the form of a syphon (fig. 5): fix tightly by means of thin wire or silk cord on the end of one limb of the tube an india-rubber blowing ball: introduce the other limb of the tube into an inverted beaker: then squeeze the ball as tightly as possible and place the beaker, mouth downwards, into a vessel containing water: the water will not enter the beaker, except the small quantity due to the compression of the air, on account of that vessel being filled with air. Remove the pressure from the ball: then, owing to the partial vacuum in the ball, some of the air in the beaker will pass into it through the tube: as the air passes out of the beaker, water will ascend into it and occupy the space the air had occupied. When all the air, or the volume sufficient to fill the ball, has passed out of the beaker, again squeeze the ball, when the air will ascend through the water into the beaker, and displace the volume of water which had entered it when the air was withdrawn.

49. The difficulty of pouring a liquid into a vessel having only one small hole arises from the impenetrability of the air, as the liquid can run into the vessel only as the air makes its escape. The following experiment illustrates that fact:—

12. Fit up an ordinary wide-mouthed bottle with funnel tube and bent tube as shown in fig. 6, immerse the outer leg of the tube in a glass containing water: then pour into the funnel water in small quantities at a time, and it will be seen that, as the water descends into the bottle, the air makes its escape through the tube, as is proved by the ascent of the bubbles of air in the water in the glass.

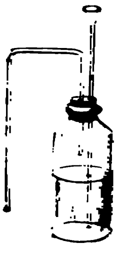

Fig. 6.

50. It not unfrequently happens that students, when first commencing the study of practical chemistry, find, to their surprise, that in performing a filtration the liquid does not pass through the filter; this is due in many cases to the leg of the funnel fitting air-tight into the mouth of the vessel into which the funnel dips: hence as the air cannot escape from the vessel, the liquid cannot enter. The following experiment is given to impress this fact on the student's mind:—

13. Fit the leg of a funnel air-tight into the neck of a bottle by means of a perforated cork (fig. 7); then place a filter in the funnel and filter;

only a small quantity of the liquid will enter the bottle, owing to the impenetrability of the air. The small quantity of the liquid which enters is due to the compression of the air by the liquid in the funnel.

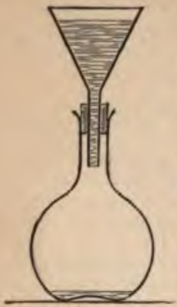

FIG. 7

51. As impenetrability and displacement are being constantly applied by the chemist in *collecting* and *storing* gases, this is the most appropriate place in the book for explaining to the student the apparatus, &c., employed for collecting gaseous bodies.

52. All my readers will have noticed the vessels—gas-holders—employed at gas-works for holding the gas; these vessels may be likened to a tumbler placed with its mouth downwards in a vessel containing water; at the first start the holder is filled with water; then as the gas passes into it the water is displaced, and this is continued until the holder is full of gas, and the gas being lighter than water the holder rises in the air: when the gas is being drawn off into the street-mains the water re-enters the holder, and it again sinks into the outer vessel.

53. All gases are collected in the laboratory by displacing some liquid or gaseous body; when the gas to be collected is much heavier or lighter than air, it is frequently collected by displacing the air in the vessel in which it has to be collected. When the gas to be collected is heavier than air, the delivery tube from the vessel in which the gas is generated must pass down to the bottom of the collecting vessel, a piece of paper or a glass plate partially covering the mouth of it, during the time it is being filled with the gas; when it is filled, the delivery tube must be removed, and the mouth of the bottle closed as speedily as possible with the stopper. When the gas is lighter than air, the extremity of the delivery tube must pass *upwards* instead of *downwards*, and the vessel in which the gas has to be collected must be inverted over the extremity of the tube, the latter passing up to the bottom of the vessel. The collecting vessels employed in both these cases are wide-mouthed glass stoppered bottles of about a quart size.

54. By the other methods for collecting gases, the vessels in which they have to be received are first filled either with the bore *straight*, and especially *equal*. It is necessary, in order to make an equal bore, to introduce the file alternately, at each end of the cork ; the hole can, of course, be made of any size by filing, and employing larger files as the size of the aperture increases.

60. After having selected a piece of hard glass tube of the proper bore and thickness, the next thing, if it does not require bending, will be to cut it into the required lengths. For this purpose, a mark is made round the tube, at the proper place, with the edge of a three-cornered file ; the tube is then held in the hand as you would a stick you were about to break ; a slight and sudden pressure is then exerted at the mark, which causes the tube to break evenly in two. The depth of the mark made with the file must vary with the thickness and diameter of the tube, and in filing very small or thin tubes, the manipulator must not press too much on the file, otherwise the tube will be broken. When a tube has to be introduced into a cork, it is necessary to round its sharp edges before doing so, else the cork will be so cut and torn that it will be rendered perfectly useless. To effect this, the tube must be heated in the flame of a gas or spirit lamp ; it must be heated gradually, so as to prevent its cracking, by holding it near the flame at first, and by degrees introducing it. As soon as the edge becomes faintly red-hot, the thing is accomplished, and the tube ought then to be removed from the flame, otherwise the opening at the extremity might contract, which, to say the least, is better avoided. If the heat of the lamp be not sufficient to round the edges it must be rounded in the blow-pipe flame. A hot tube must not be laid on any cold surface, and it must be allowed to cool before inserting it into a cork. The perforation in the cork must, if anything, be a trifle smaller than the diameter of the tube which has to be inserted, in order that the connection may be perfectly air-tight. The best plan of fitting the tube into the cork is to screw the tube, holding both it and the cork firmly, the first close to the extremity which has to be inserted; the tube may be greased very slightly at the part to be introduced.

61. Gas jets are sometimes required in the combustion of gases ; they are made in the following way :—a hard glass tube of the proper diameter and thickness is heated in a spirit or gas lamp, at such a distance from one extremity that there i sufficient left as a handle for the manipulator. and also sufficiently remote from the part to be heated that it does n become insupportably hot. The part to be heated is to b brought, not into the flame, but into the current of hot a which passes off in the same direction from it, and the tube i to be turned, so as to become heated all round, and also move a little to the right and left, that the temperature of the neighbouring parts may be raised. After a few seconds, when the glass has become hot, it should be brought towards the point of the flame, and ultimately into it, being *turned round* all the time, and also moved laterally, though not to the same extent as before. When sufficiently soft it is drawn out into a capillary, as shown in fig. 13. The two ends are then separated at about

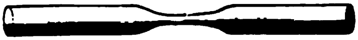

FIG. 13.

the centre of the portion drawn out. The elongated extremity of the end to be used is filed down until the orifice is of sufficient size; it is then rounded at the edges in the flame of the lamp.

62. The flame of a gas lamp, using a fish-tail burner, is sufficient for bending ordinary glass tubing without the aid of the blowpipe. The tube must be heated about an inch in length uniformly; it must be supported by the hands on each side of the part to be heated; it must be gradually introduced into the flame, in the way described in the preceding paragraph, and turned round continually during the process of heating, as there stated. Both ends must be turned alike, and in such a way that the heated part, when soft, shall retain its cylindrical shape, being neither twisted, nor bent, nor elongated, nor thrust up. When the glass has become sufficiently soft, remove it from the flame, and gently bend it to the desired shape. The bend ought not to become flat at the convex portion, o in folds at the concave; to avoid these imperfections, the glas ought not to be made too soft, it ought to be heated uniforml to the length of about an inch, and the bend ought to be mad very round. The side of the tube last exposed to the flam will be slightly hotter, and therefore softer, than that which i

opposite to it. This hotter side should form the concave side of the bent tube.

63. Vulcanised india-rubber tubes are employed for joining glass tubes; they can now be purchased of any size, and therefore the chemist has no need to make them. In connecting glass tubes by means of the rubber tubing, the ends of the glass tubes should not be brought quite close together inside the rubber connector, but should be left a very slight distance apart, in order to allow of a certain degree of flexibility. If the vulcanised india-rubber tubing does not fit the glass tube air-tight, it must be tied on with some silk cord.

64. The manipulator having carefully attended to the preceding instructions, will experience very little difficulty in fitting up the different apparatus now to be described.

65. **Apparatus for generating gases, heat not being required.**—A bottle, with two or three separate openings, called a **Woulfe's Bottle,** is a very convenient vessel for generating gases when heat is not required. It need only have two openings ; one for the funnel tube, which serves for the introduction of liquid, and which must pass down nearly to the bottom of the vessel, being covered by the liquid in the vessel to the extent, at least, of half an inch ; the other, for the conveying of the gas to the receiver ; this is called the delivery tube ; it need only project into the bottle a trifle beyond the cork. The delivery tubes are of various shapes. If the bottle is provided with three openings, the middle one can be closed with a plain cork, and be used for introducing any solid substance into the bottle. Wide-mouthed bottles may be used in the place of Woulfe's bottles ; the only objection to them is the difficulty of making them air-tight, as large corks, free from pores, can seldom be obtained. If any of the corks be porous, it will be advisable, after the cork has been fitted into the apparatus, to paint the top of the cork with a solution of sealing-wax in spirit of wine. But if the tube does not fit the perforation in the cork perfectly air-tight, a paste, made with linseed meal and water, or better still, with lime-water, ought to be used for stopping the leakage. Gases should not be generated too rapidly ; if an acid be employed for disengaging the gas, it ought to be added by degrees, commencing with a little, and

when the gas begins to be disengaged *slowly*, to add a little more, &c. The delivery tubes are, as we have said, of various shapes; fig. 14 represents a Woulfe's bottle fitted up in the way just described, the delivery tube being made of the shape required for collecting the gas at the pneumatic trough, or in the gasholder. The other (fig. 15) is fitted up for collecting the gas by displacement, or for passing it through a solution. Sometimes it is necessary to wash the gas, which is done in the following way; a wide-mouthed bottle, capable of holding about half a pint, is fitted with two tubes, as shown in fig. 16, one of which, *a*, should be about half an inch in diameter, straight, and of sufficient length to reach nearly to the bottom of the bottle, the water in which ought to cover the tube to the depth of half an inch; the other (the delivery tube) should only just pass through the cork, and be bent according to the mode of collecting the gas. The delivery tube of the vessel, in which the gas is generated, is made to pass down the tube *a*, and project a little beyond it in the bottle; the gas will then rise up through the liquid and pass off by the delivery tube *b* of the washing-bottle.

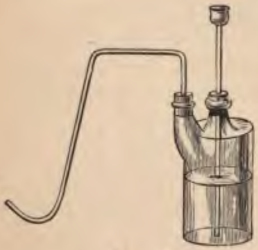

FIG. 14.

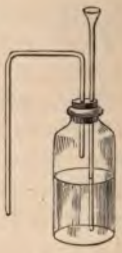

FIG. 15.

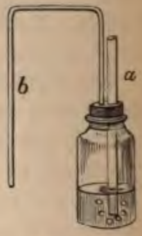

FIG. 16.

66. Apparatus used in the disengagement of gases, heat being required.—Flasks, test-tubes, and retorts are the ordinary kinds of apparatus used for this purpose. Figures 16 and 17 represent the first two, with their different fittings and connections. The first are used in preference to retorts on account of their cheapness, and the second are the most convenient when only a small quantity of gas is required. The materials used in the production of a gas ought only to occupy, at the most, one third of the space in the generating vessel, whether heat be employed or not. When the materials for the production of the gas consist of a liquid and a powder, the liquid ought to be put into the vessel first, and after the powder has been added, they ought to be well mixed by shaking the vessel before heat is applied.

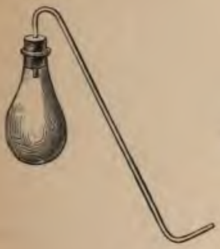

FIG. 17.

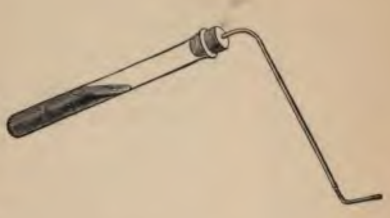

FIG. 18.

67. Heat has to be employed in a great many chemical operations ; gas is the most convenient and economical fuel; it requires to be mixed with a certain quantity of air in the tube of the lamp before being burnt, so as to prevent the deposit of any soot on the vessels. A cheap and most convenient lamp is that called Bunsen's gas-burner ; it consists (fig. 19) of a brass tube, *a*, *b* ; these tubes are made of different diameters ; air enters by holes at *c*, the amount being regulated by means of a small slide, which closes them wholly or partially as may be required. The gas enters at *d*; when a more diffuse flame is required an iron rose (fig. 20) is screwed on to the top, the rim of which is perforated with a number of small holes. When gas cannot be obtained, wood naphtha or methylated spirit is

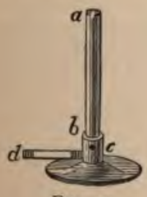

FIG. 19.

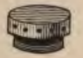

FIG. 20.

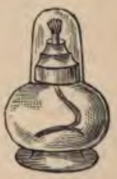

FIG. 21.

the fuel usually employed ; where a great amount of heat is not required it is burnt in the lamp (fig. 21). When the lamp is finished with, the cap ought to be placed over the wick in order to prevent the loss of the liquid by evaporation. The cap ought not to be used as an extinguisher, because the air in the cap becomes, if it is so used, so rarefied that it is afterwards difficult to remove it, but the flame ought to be *blown out* before the cap is replaced. The heat ought to be applied gradually for two reasons: (1) in order to prevent the breaking of the vessel, which is almost sure to be the case if the heat be applied suddenly ; (2) to prevent a tumultuous disengagement of the gas ; when the evolution of the gas diminishes in rapidity, the heat ought then to be increased. A retort or flask is supported on one of the rings of a retort-stand (fig. 22) ; a test-tube is held over the lamp by the fingers.

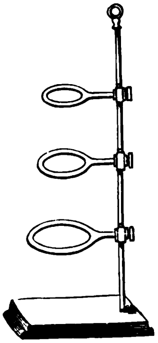

FIG. 22

Examination Questions.

14. Define the meaning of the word 'matter' as applied to the material world.

15. Describe the operation called 'welding.'

16. Define the meaning of the terms 'particle' and 'molecule.'

17. What precautions are necessary in getting out of an omnibus in motion, so as to avoid falling to the ground? assign a reason for what you state.

18. State the difference between specific and general properties.

19. What properties are bodies said to possess when they admit of being bent without breaking, and retain the form they have been made to assume?

20. Define the meaning of the term inertia as it has been employed in the text.

21. What are the external signs by which the existence of matter in its various forms is made evident to our senses?

22. Give a sketch of the apparatus you would employ, and in its proper position, if you had to collect a gas lighter than air ; also the apparatus for a gas heavier than air ; each of the gases being collected by displacing the air in the receiving vessel.

23. When the term 'hard' is applied to solid bodies, what does it signify?

24. What is meant by the terms 'soft' and 'brittle' when applied to solid bodies?

25. What is meant by the term 'elasticity' when applied to bodies?

CHAPTER II.

MOLECULAR ATTRACTION—GENERAL EFFECTS OF HEAT—MAXIMUM DENSITY OF WATER—HYPOTHESIS AS TO THE NATURE OF HEAT—DIFFERENCE BETWEEN HEAT AND TEMPERATURE—MEASUREMENT OF TEMPERATURES—EXERCISES—SPECIFIC HEAT, AND THE METHODS FOR ITS DETERMINATION—EXERCISES—LATENT HEAT—QUESTIONS—DYNAMICAL THEORY OF HEAT—QUESTIONS.

68. THE molecules of all substances are under the influence, as has been previously stated, of two contrary forces; one of which tends to bring them together, the other tends to separate them from each other. According to the manner in which the molecular attraction—the attractive force—operates, it is designated by the terms *cohesion*, *adhesion*, or *chemical affinity*; cohesion has already been explained; adhesion and affinity will be explained subsequently; the opposing force has here to be considered.

69. The repulsive force exerted between the molecules of substances is developed by the action of heat; hence one of the effects of heat upon bodies is to cause them to expand, in other words, to increase in volume; and also to change the state of aggregation of solids and liquids—that is, it causes solids to pass into the liquid state, and liquids to pass into the gaseous state. When a solid body, for example, is exposed to the influence of heat, it expands according to *three* dimensions, viz. 1. In length, called its *linear* expansion. 2. In surface, called its *superficial* expansion. 3. In volume or bulk, called its *cubical* expansion: and when the solid is perfectly homogeneous, it expands uniformly throughout its entire volume. The increase of surface is twice, the increase of volume is thrice, the increase of the linear expansion; and one of these expansions, it may be observed, never takes place without the other. That solids increase in volume by increase in temperature is easily shown: an iron or brass ball, which just goes through a ring when cold, will not fall through it when heated; but if left on the ring will, on cooling, fall through it. This experiment not only proves that solids *expand* on being heated, but it likewise proves that they *contract* on cooling. There is a very common and useful application of the expansion and contraction of solids in the hooping of the wheels of carts, carriages, &c.; the hoop or band of iron is purposely made less than the wheel itself, so that it cannot go on when cold; but when it is heated it expands sufficiently to be placed on the wheel, and when in its proper position cold water is poured upon it, which causes it instantly to contract and bind the wheel tightly together.

70. Solids dilate pretty regularly, at temperatures between the freezing and boiling points of water; consequently, equal increments of heat correspond with equal increments of volume. At higher temperatures, the expansion proceeds at a greater rate than the successive additions of heat. The degree of expansibility varies, of course, in different substances.

71. One or two solid bodies are known which, at least within certain limits of temperature, form exceptions to the general rule of expansion by heat, for they contract as the temperature is raised up to a certain degree of temperature. An alloy consisting of 2 parts bismuth, 1 part tin, and 1 part lead *expands* when heated from 0° to 44° C.; when still further heated, it *contracts*, so that at 56° its density is the same as it was at 0°, and at 69° still greater; beyond this temperature, expansion again takes place; at 87·5° the alloy has once more the same density as at 0°, and at 94° the same as at 44°.

72. Liquids, like solids and gases, expand when they are heated and contract when they are cooled; they expand proportionately much more rapidly than solids; they differ also in expansibility to a much greater extent, the most volatile being generally the most expansible. The rate of dilatation of liquids is not generally uniform, like that of solids and gases, it increases with increase of temperature; and as the liquid approaches solidification on the one hand, and the point of vaporisation on the other, other irregularities in its contraction and expansion take place.

73. Water is a most remarkable *exception* to the law of expansion by heat and contraction by cold; it contracts on cooling like other liquids until it reaches the temperature of 4° C. (39·2° F.); at this temperature it attains its greatest density, for on further cooling it begins to expand, and just before it begins to freeze, its volume at that temperature and that at 9° C. are equal : if then the temperature of water either rises above or sinks below 4° C. the water *expands*. On freezing the expansion is so great and sudden that the force exerted is enormous ; thick iron shells, filled with water and their fuse-holes securely plugged, when they have been exposed to the severe cold of a Canadian winter night, have been found split open in the morning. In our comparatively mild winters we are made aware of this great expansive force in the bursting of water-pipes and breaking of vessels by the freezing of the water they contain ; the expansion is nearly $\frac{1}{11}$ the volume of the water. The student must, however, note the distinction between the expansion on freezing, and the expansion as the water cools from 4° C. to the point it begins to freeze ; it is this latter expansion that is anomalous, for no other substances besides water expand on solidifying.

74. The freezing of water in the joints and crevices of rocks acts as a disintegrator and pulveriser even of the hardest of them, thus facilitating the chemical action of the carbon dioxide in the air in converting rocks into fruitful soil. This anomalous expansion makes water perform other important functions in the economy of Nature ; were it not for its expansion after a certain temperature is reached, lakes and rivers would become solid masses of ice in winter, which the summer heat would not be sufficient to melt, and all aquatic life in these waters would necessarily perish. This is prevented, owing to the cooling taking place at the surface ; the surface water becoming heavier than that underneath sinks and is replaced by less dense water ; this replacement of the surface water by that underneath continues until the whole mass acquires a temperature of 4° C. ; when this temperature has been reached the surface water ceases to sink, owing to its becoming of less density than the water underneath ; and the lighter water on the top protects that underneath from a further reduction of temperature owing to its bad heat-conducting power ; and when it solidifies by further cooling, the ice formed remains on the surface, because its density is less than that of the water underneath. When the surface is frozen, the tempera-

ture of the deep water remains at 4° C.; aquatic animals are thus kept in a temperature high enough for their existence.

75. Several methods have been devised for determining the maximum density of water; the one we describe is that devised by Hope; it consists of a deep glass tube-shaped vessel (for lecture experiments it is frequently made of tin as it renders it less costly) having two lateral openings, one near the top and the other near the bottom; a thermometer is placed in each of these openings (fig. 23). The tube is filled with water, and its middle is surrounded with a freezing mixture. 'The following phenomena will then be observed.

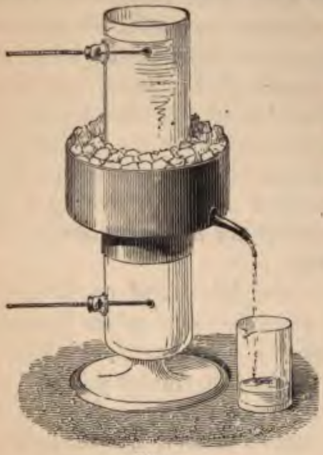

FIG. 23.

76. 'The lower thermometer descends steadily to 4° C., and there remains stationary. The upper thermometer at first undergoes very little change, but when the lower one has reached the fixed temperature, the upper one begins to fall, reaches the temperature of zero, and, finally, the water at the surface freezes, if the action of the frigorific mixture continues for a sufficiently long time. These facts admit of a very simple explanation.

'As the water in the middle portion of the tube grows colder, its density increases, and it sinks to the bottom. This process goes on till all the water in the lower part of the vessel has attained the temperature of 4°. But when all the water from the centre to the bottom has attained this temperature, any further cooling of the water in the centre fails to produce motion until needles of ice are formed. These being specifically lighter than water rise to the surface, and thus produce a circulation which causes the water near the surface to freeze, while that near the bottom remains at the temperature of 4°' (Deschanel).

77. Salts dissolved in water lower its point of maximum density, common salt (sodium chloride) especially; in sea-water, which contains an abundance of common salt, the maximum density occurs at 13·68° C. (25·38° F.), a temperature below its freezing-point.

78. The only body in the liquid state whose variation of volume through a considerable range of the thermometric scale is found to be exactly proportional to its change of temperature, is mercury.

79. From what has been stated with regard to the expansion of solids by heat the student will at once see that the increase of expansion of any liquid, in a vessel, is not its *real* or *absolute* expansion, but only its *apparent* expansion; in order to arrive at the absolute expansion of a liquid, the changes in the capacity of the vessel at the same temperature which contains the liquid must be determined; *the apparent expansion of a liquid is therefore equal to its absolute expansion for the same interval of temperature diminished by the corresponding expansion of the containing vessel.*

80. In the expansions of liquids, the cubical expansion is alone considered.

81. All gases, on account of the absence of any cohesiveness between their particles, *expand alike*, *the pressure being the same for equal* additions of heat; the rate of expansion being equal and uniform at all degrees of temperature; they expand rather more than one-third their volume on being raised from the temperature at which ice melts to that at which water boils at the *ordinary atmospheric pressure.* The expansion of gases is more fully treated of in Chap. III.

82. The two prevailing opinions with regard to the nature of heat, are the *theory of emission*, and the theory of *undulation*.

83. According to the first theory, heat is a subtle imponderable fluid, which surrounds the molecules of bodies and which can pass from one body to another. According to the other theory, which is the one now generally accepted, the heat of a substance is caused by the oscillatory motions both of the particles of bodies, and of a very subtle medium which is considered to pervade all space, and which is capable of transmitting a vibratory motion with great velocity. 'A rapid vibratory motion of this ether produces heat, just as sound is produced by a vibratory motion of atmospheric air, and the transference of heat from one body to another is effected by the intervention of this ether.' Heat, according to this view, is not a *substance*, but a *condition of matter*, and a condition which can be transferred from one body to another; it is, in short, looked upon as a *form of motion*.

84. The general effects of heat we have already noticed; here we will make one remark before going further into the subject:—whenever, as in the case of heat, we have to speak of any agent whose nature is unknown, the expressions and allusions we employ in regard to it must necessarily possess a degree of vagueness; and to those students who require exact definitions, they must appear unsatisfactory. This vagueness is further increased if different views have prevailed at different times as to the nature of the agent, because the terms which have been employed under one theory are frequently retained when other and totally different views prevail. These few general observations on the terms employed we have thought it necessary to make, in order to show the student that the terms which may appear vague and unsatisfactory to him are equally so to all who study the subject.

85. The term heat we apply to the agent; the term temperature we apply to that portion of heat in a body which we can perceive by the senses. The temperature of a body can be increased or diminished. If two bodies of different temperatures are brought together, the one containing the most sensible heat—in other words, having the highest temperature—will lose heat, which the body with the lower temperature will absorb. This imparting of heat by the one body, and the receiving of it by the other, will only cease when they have both arrived at the same temperature. If, therefore, we bring the hand, or any other part of our body, in contact with a substance having a higher temperature, heat is imparted by that substance to the hand, and we experience the sensation we term heat. If, on the contrary, the hand is warmer than the substance, it imparts some of its heat to the substance, and we then experience the sensation we term cold. The student will observe, from these illustrations, that heat and cold are merely relative terms.

86. Our sense of touch cannot be employed to ascertain the correct temperature of substances ; for that will only tell us whether the substance has a higher or a lower temperature at the time than that part of our body which comes in contact with it. The same body might even be made to be feel both hot and cold *at the same moment*, if we made one hand hotter and the other colder than the body, before we brought them in contact with it. For the same reason, we pronounce a cave cold in summer, hot in winter, although the temperature of the cave may have remained the same. To measure the correct temperature of bodies, we have recourse to the physical action of heat on bodies. These actions are of various kinds. Expansions and contractions have been adopted, as being the easiest to observe. But heat also produces electrical phenomena in bodies, on which very delicate methods of observing temperatures have been based.

87. Instruments termed thermometers (from θερμὸς, hot, μέτρον, a measure) are employed to measure temperatures up to about 600° F. Instruments termed pyrometers (from πῦρ, fire), are employed to measure temperatures higher than 600° F.

88. Although bodies in each of the three states of aggregation expand on being heated and contract on cooling, they are not equally suitable for thermoscopic agents. Solids, as a rule, are not suitable, as their expansion is comparatively slight ; their increase, therefore, in volume for an increase of temperature cannot be easily or accurately observed. Air is much more delicate in its indications of temperature than liquids, and its expansion being uniform at all degrees of temperature (par. 81) would render it the best thermometric agent, were it not that a difference in atmospheric pressure, independent altogether of temperature, causes an alteration in its volume. Elaborate and special contrivances, consequently, are required to render it a reliable thermoscopic agent ; its use is therefore limited; it is chiefly used in ascertaining the specific gravities of volatile substances, and other delicate experiments. Liquids, as a rule, are best suited for measuring changes of temperature, and the liquids used for this purpose are mercury and alcohol.

89. Mercury is the best thermoscopic agent for general purposes, for it has the widest range between its solidifying and boiling-point of any liquid; it also more nearly resembles gases in uniformity of expansion for equal increments of heat, at ordinary temperatures, than any other liquid (par. 78). It is likewise easily purified; it is also a very good conductor of heat, and therefore becomes almost immediately of the same temperature as the substance it comes in contact with; and, owing to its calorific capacity being very small, the heat it abstracts from the heated body it is in contact with causes such a slight change in the temperature of the heated body, that this slight lowering of temperature may be generally neglected, and, owing to its small calorific capacity, it is rapidly heated and cooled.

90. Alcohol is employed to measure very low temperatures, as it does not solidify at the greatest known cold.

91. We must refer the student to the general works on chemistry and physics for descriptions of the differential thermometer, the thermo-electric multiplier, the self-registering thermometer, the air-thermometer, the pyrometer,—in fact, all the varieties of heat-measurers but the mercurial thermometer, which all my readers will have seen, at least in its most inaccurate form, viz. that used for common purposes, and which is provided with a boxwood scale.

92. We shall not enter into the minutiæ for constructing a mercurial thermometer; but we will very briefly state the essentials and the mode for testing its accuracy, as it is a very important and indispensable instrument in most physical and chemical researches. The first essential in the manufacture is to select a capillary glass tube of uniform bore throughout its entire length, so that *equal lengths* of the tube will contain *equal volumes* of mercury. To ascertain this, a small drop of mercury is introduced into the tube, sufficient to fill about one inch of the tube. The mercury is gradually moved from one end of the tube to the other, and its length is measured in each successive position. The length, of course, should be the same in every part of the tube; if it be not, the tube must be rejected. Having selected the tube, the next thing to be done is to blow a bulb upon one end of it. The ordinary shape of the bulb is spherical, but an elongated tubular-shaped bulb is more convenient in most chemical investigations, and as it exposes a much larger surface for the conduction of heat, in proportion to its capacity, it is consequently more sensitive. The thickness of the glass is an important point; if too thin it is liable to be altered in shape by the pressure of the atmosphere, and thus the graduation would be rendered inaccurate; and if it be too thick, it will render the instrument insensitive. The instrument is now ready to be filled with mercury, if the interior of the tube is thoroughly clean and dry. 'The following is the most perfect mode of filling the thermometer with mercury:—A sufficient length of the tube, *a b*, is marked off to form the stem of the thermometer, and a second bulb, rather larger than the first is roughly blown at *b*. When the bulb is half-cooled, the open end of the tube is held in pure and well-boiled mercury, which gradually rises into the first bulb. By inverting the position of the tube, and alternately heating and cooling the lower bulb (*a*), sufficient mercury may be made to pass from *b* into *a*, almost to fill *a*, a little remaining over in *b*. The whole tube is now held by the handle, shown in fig. 24, over a charcoal fire or a series of powerful gas-burners, so as to be strongly and uniformly heated throughout its whole length. The mercury boils easily and its vapour drives out every trace of air or moisture. When this is thoroughly accomplished, and the mercury yet boiling, the open end of the tube, which must not be very hot, is thrust against a piece of sealing-wax, and then at once withdrawn from the heat. As the tube cools, a little wax is drawn in, and, in time becoming solid, prevents the entrance of any air. When quite cool, the mercury perfectly fills the whole of the lower bulb and tube. Let the tube be now inclined, and the lower bulb raised to a temperature rather higher than the thermometer is intended ever to indicate. A part of the mercury is thus expelled, and as the column begins to retreat again on cooling, the tube is to be melted and hermetically sealed just below the bulb *b*, which is then drawn off. The perfection of the vacuum thus produced within part of the

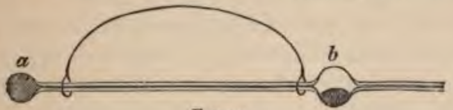

Fig. 24.

tube may be known by the mercury running freely up and down, especially if the tube be not very fine. This process is more simply but less perfectly performed by some makers, who omit the second bulb, and place the mercury in a little funnel of glass or paper affixed to the open end of the tube.' At the time of sealing the tube, it contains, or at least should only contain, mercury and vapour of mercury, all trace of air and moisture having been expelled by boiling the mercury in the tube.

93. The thermometer has, of course, to be graduated, that is, provided with a scale to which variations of temperature can be referred; and in order to be able to compare one thermometer with another, it is necessary, in the graduation of thermometers, to select *two standard temperatures as fixed points*; the temperatures which have been universally chosen, are those of *melting ice* and *boiling water*. The thermometer ought not to be graduated immediately after it is sealed, as the bulb contracts, and does not attain its fixed dimensions until some months after the sealing; consequently, if the fixed points were determined before the bulb arrived at its fixed capacity, the graduations on the scale would indicate a temperature higher than the true one from a quarter to half a degree.

94. Owing to a frequent rise in the melting-point after graduation, which is generally supposed to be due to atmospheric pressure, the question whether the boiling-point or the freezing-point ought first to be determined is a matter not settled. Most scientific men consider that the freezing-point ought to be determined first; others, on the contrary, think that the boiling-point ought to be taken before the freezing-point; whichever plan be adopted, a considerably long interval ought to elapse between the taking of the two points.

95. To determine the freezing-point the thermometer is placed upright in a vessel containing snow or finely pounded ice, and which has a hole at the bottom to let the water drain away as the ice melts. The whole of the mercury, both in the tube and in the bulb, must be surrounded with ice, and must be left in it for half an hour; the ice may then be removed just sufficiently to allow a mark to be made at the top of the mercurial column; this mark of course represents the point on the scale at which ice melts; the instrument is now removed.

96. 'The boiling-point is ascertained by placing the thermometer vertically in a cylindrical metal vessel, so that the bulb shall be close *above* the surface of distilled water in rapid ebullition, the whole of the mercurial column being likewise raised to the same temperature as the bulb by the rising vapour. The water having been thoroughly boiled for a length of time, and the mercury standing quite stationary, a mark is made at the top of the mercurial column, and this must be done while the boiling continues uninterruptedly. As the boiling-point of water depends entirely on the pressure of the air at the time, the pressure, by means of the barometer must be determined. If it stand at 760 millimetres (or 30 English inches), the temperature of boiling water is exactly 100° C. (212° F.). If not, the temperature must be calculated, allowing a change of ± 1° C. for every divergence of pressure equal to ± 26·7 millimetres (or 1·71° F. for every inch of pressure).'

97. The two fixed points having been marked on the tube, the distance between them is divided into equal parts, termed degrees. The degrees are designated by a small cipher placed a little above on the right of the number which marks the temperature; and, to indicate temperatures below zero, the minus sign is placed before the figures thus:—15° signifies 15° below zero. In *accurate* thermometers, the scale is marked on the glass itself, which is accomplished by covering the glass with a thin coating of beeswax, then marking the divisions and numbers with a steel point, and subsequently exposing the thermometer for a short time to the action of hydrofluoric acid gas: where the wax has been removed by the steel point the glass is etched by the gas.

98. There are unfortunately three different thermometric scales, viz., the Centigrade, or Celsius (the latter the name of the person who invented the scale), the Fahrenheit scale (this scale originated with an instrument-maker at Dantzic), and the Réaumur scale. The Centigrade scale is adopted on the Continent, especially in France, and it is the one most generally employed in scientific investigations. The Fahrenheit scale is universally employed in England for ordinary purposes, and frequently in purely scientific investigations; it is also employed in America and Holland. The Réaumur scale is employed in Russia and Germany.

at a convenient distance above the first, and dividing the intervening length into an arbitrary number of equal parts. Thus all is arbitrary ; and there is no peculiar virtue in the two points which have been chosen, other than that they can be easily determined with accuracy, and include between them the range of temperature with which we are usually most concerned.'

105. When a thermometer is introduced into melting ice, the level of the mercury in the tube should stand exactly at the point of the scale which indicates this temperature ; and when it is suspended in the steam of boiling water, the level of the mercury should stand, at the *standard pressure*, exactly at the point of the scale which indicates the temperature of boiling water. To ascertain whether the value of the degrees is uniform throughout the entire length of the tube, a little column of mercury must be detached from the rest of the metal, the tube must then be inclined so as to make the small column pass from one part of the tube to another part, and we must ascertain by measurement whether it occupies the same length in the tube in each position. To ascertain whether the air has been perfectly expelled before sealing the tube, the instrument must be inverted, so that the bulb is uppermost; if the mercury descends with ease *quite* to the end of the tube, we may be satisfied that air is absent in that part of the instrument. But there may be some in the bulb—to ascertain this, we must observe whether the bubble which is in the bulb during the time it is inverted entirely disappears when we restore the instrument to its proper position ; if it does, the thermometer is perfect as regards the entire absence of air.

106. 'Sometimes a thermometer, when inverted, does not allow the mercury to descend, and inexperienced persons are liable to imagine that air must necessarily be present ; but the effect alluded to is frequently caused by the adhesion between the mercury and the glass ; if, however, we warm the bulb, so as to cause the mercury to ascend in the tube, and then invert the instrument, the metal should descend ; and if no moderate tapping or swinging will make it do so, we may safely conclude that air forms the impediment.'

107. As it is frequently necessary to translate the figures on the Centigrade scale into those of Fahrenheit, and *vice* *versâ*, tables are constructed by which the degrees on the one scale can be converted into those of the other at a glance. But it is necessary to know how to convert the one into the other by calculation, which is easily accomplished; 100° C. = 180° F., or 5° C. = 9° F.; hence, 1° F. $= \frac{9}{5}{}^\circ$ C. $+ 32$: the 32 is added because the zero of Fahrenheit is 32 Fahrenheit degrees below the zero of the Centigrade. The conversion is easily effected by the use of the following formulæ:—

Fahrenheit to Centigrade	$\frac{5}{9}(\text{F.}^\circ - 32) = \text{C.}^\circ$
Centigrade to Fahrenheit	$\frac{9}{5}\,\text{C.}^\circ + 32 = \text{F.}^\circ$
Réaumur to Fahrenheit	$\frac{9}{4}\,\text{R.}^\circ + 32 = \text{F.}^\circ$
Fahrenheit to Réaumur	$\frac{4}{9}(\text{F.}^\circ - 32) = \text{R.}^\circ$

Exercises on the conversion of one thermometric degree into another of a different scale.

26. Reduce 60° F. to the degrees of C. and R.; 60° of R. to F.°; 50° C. to F.°; 20° F. to C.°; −2·2° F. to C.°; −2° C. to F.°; −5·8° F. to C.°; 5° C. to F.°; −24° C. to F.°; −36° C. to F.°

108. If equal quantities of the *same* liquid, at *different* temperatures, are mixed together, the temperature of the mixture will be the *mean* of the two; thus, if one pound of water at 60° is mixed with the same quantity at 212°, the temperature of the mixed mass will be 136°. But if equal weights of *different* liquids are mixed together, the temperature of the mixture will not be the *mean* of the two; thus, if one pound of mercury at 40° is mixed with the same quantity of water at 156°, the temperature of the mixture will be 152·3°. Therefore, whilst the temperature of the water has only been depressed 3·7°, enough heat has been evolved to raise the temperature of the mercury 112·3°. If we substitute volumes for weights, we find that water requires double the quantity of heat that an equal volume of mercury requires, to raise it to a certain number of degrees; thus, if two glass vessels, one containing water, and the other an *equal volume* of mercury, be immersed at the same time in a hot water bath, it will be found that the mercury becomes heated to the temperature of the water in the bath, in *half the time* that the water in the glass vessel requires to gain that temperature; and if the vessels containing the mercury and water be removed from the bath, after the two liquids have arrived at the same temperature, the mercury, it will be found, will cool *twice as rapidly* as the water. To take another example; if one pound of iron filings at 68° be placed in one pound of water at 32°, the temperature of the mixture will be 36°; consequently, the quantity of heat, the loss of which lowers the temperature of the iron 32°, increases the temperature of the water only 4°. It is evident, from these examples, that equal weights and equal volumes of *different* substances require the addition of different quantities of heat to produce the same alteration in their temperature; hence, different bodies at the same temperature, and whose volumes or weights are equal, contain *unequal* quantities of heat.

109. From what has just been stated, the student will clearly see that the thermometer only measures the temperature of bodies, and not the amount of heat they contain; for different bodies, at the same temperature, and whose volumes or weights are equal, contain unequal quantities of heat. Unequal quantities, by volume or weight, of the same body, may also indicate the same temperature, though the larger quantity must necessarily contain more heat than the lesser quantity; for instance, the water in a wineglass will have the same temperature as the water in a cistern, if filled from the same source. In one condition only the thermometer becomes a direct measure of the amount of heat, and that is in the case of the same weight of the same substance. Thus if we take the one pound of water, it is true that, if a given amount of heat will raise its temperature 1°, twice the amount will raise its temperature 2°; here, then, we have a unit for measuring the amounts of heat.

110. The absolute quantity of heat which a body gains or loses cannot be measured; but we can ascertain how much one body gains or loses compared with another, either by the process of mixing described in the examples we have given, or by other methods we shall presently describe; we ascertain by these methods the relative quantities of heat which equal volumes or weights of different substances require, to raise them to a given temperature. Thus we have already noticed, that if one pound of mercury at 40° F. is mixed with one pound of water at 156° F., the mixed mass will have a temperature of 152·3° F. Consequently the same weight of water requires, according to these numbers, rather more than thirty times as much heat as mercury requires to produce the same elevation of temperature; as mercury requires a less amount of heat than water to raise it a given number of degrees, it is said to have *a less capacity* for *heat* than water. The capacities for heat of all solid and liquid bodies are expressed in numbers in relation to that of water; these numbers are termed the *specific heats* of bodies.

III. Quantity of heat must, like all other quantities, be measured in the terms of *some unit*; pure water is taken as the standard substance in measuring quantities of heat; but several different units of *heat*, and units of *mass*, are employed. In England, for practical purposes, the unit *mass* is one pound avoirdupois of water, and the unit of heat is one degree Fahrenheit, or Centigrade; but the temperature of the water from which the standard is taken varies; some adopt, as the *thermal* unit, the quantity of heat required to raise 1 lb. of water one degree F. between 55° and 60°; others adopt as the thermal unit the quantity of heat necessary to raise that quantity of water from 0° to 1° C. Whichever of the two thermometric scales is employed, the unit of heat is termed the pound-degree. In France and in all countries, for scientific purposes, the unit of mass employed is either one kilogramme, or one gramme, of water; and the unit of heat one degree Centigrade, and from 0° to 1° C.; when the gramme is employed it is termed the gramme degree. We thus learn, that by the term specific heat, or calorific or thermal capacity of a substance, is meant *the quantity of heat which is required for raising its temperature through one degree, compared with the quantity of heat which an equal weight of water would require to be raised through the same degree.* Thus, if we take five vessels, containing respectively one pound of water, one pound of sulphur, one pound of iron, one pound of silver, one pound of mercury, and expose them all to such a source of heat that equal amounts must enter each vessel during the same time—if, now, we observe the thermometers which have been placed in these vessels, we shall find, when the temperature of the water has risen one degree, and consequently, when one unit of heat has entered each vessel, that the temperatures of the other substances have increased by the number of degrees given in the second column of the following table. By the principle just established, it follows that, if one unit of heat will raise the temperature of one pound of mercury 30°, it will only require one-thirtieth as much, or 0·033 of a unit of heat, to raise the temperature of the same weight one degree. In like manner, the fractional parts of a unit of heat required to raise the temperatures of one pound of each of the other substances one degree can be easily calculated, and are given in the third column of the table ; these fractional parts are termed the specific heats of the substances :—

	Temperature	Unit of heat (Specific heat)
Water	1·0°	1·000
Sulphur	4·9	0·203
Iron	8·8	0·114
Silver	17·5	0·057
Mercury	30·0	0·033

Water, then, at the same temperature, contains 4·9 times as much heat as the same weight of sulphur, 8·8 times as much as the same weight of iron, 17·5 times as much as the same weight of silver, and 30 times as much as the same weight of mercury ; and, in like manner, we should find that at the same temperature, and for equal weights, water contains more heat than any solid or other liquid known. Hence the specific heat of solid or liquid substances is always expressed by fractions.

112. 'Three methods have been employed for determining the specific heat of solids and liquids—the method of the melting of ice, the method of mixtures and that of cooling. In the latter, the specific heat of a body is determined by the time it takes to cool through a certain temperature. Before describing these methods, it is important to explain how we measure the quantity of heat absorbed by a body of which the mass and the specific heat are given, when the temperature is elevated a certain number of degrees.'

113. *Measure of the sensible heat absorbed or lost by a body when raised or lowered from one temperature to another.* Let the unit of heat be that amount of heat which would raise the temperature of a pound of water from 0° to 1° C. If H be the specific heat of any substance, it follows from the definition of specific heat we have given, that H will represent the quantity of heat required to raise the temperature of a pound weight of such substance 1° ; 2 H the quantity required to raise it 2° ; and so on.

114. If the temperature of a pound weight of a substance, whose specific heat is H, be raised through any number of degrees, say D, the amount of heat required will be measured by $H \times D$. A body whose specific heat is H, weighing W pounds will therefore require a quantity of heat $W \times H \times D$ to raise it through D degrees of temperature.

This is usually expressed thus :

$$Q = W\ H\ D.$$

Example :—How much heat is required to raise the temperature of 167 pounds of iron from 25° to 35°, the specific heat of iron being 0·114? Here the number of degrees through which the body is to be raised is 10°, since $35° - 25° = 10°$; 167 must therefore be multiplied by 0·114 and by 10. The result is :

$$Q = WHD = 167 \times 0{\cdot}114 \times 10 = 190{\cdot}38° \text{ of heat.}$$

115. The above formula will also give the quantity (or degrees) of heat given out by a body of weight W, and specific heat H, in cooling down through D degrees of temperature. To illustrate it we will give the previous example slightly altered :—

How much heat is given out on the temperature of 167 lbs. of iron being reduced from 45° to 35°? Here the number of degrees the metal is reduced in temperature is 10° (since $45° - 35° = 10°$); 167 must therefore be multiplied by 0·114 and by 10. The result is :—

$$Q = WHD = 167 \times 0{\cdot}114 \times 10 = 190{\cdot}38° \text{ of heat.}$$

116. *Method of the fusion of ice.*—Black employed a block of ice in which a cavity was formed, and which was provided with a cover of ice, as a *calorimeter*. The substance whose specific heat had to be determined was heated to a certain temperature, and then placed in the cavity, and covered with the ice-cover. When the body had become cooled to zero, the cavity was opened, and both the substance and the cavity were wiped perfectly dry with a cloth which had been previously weighed. The cloth was subsequently weighed; the increase in weight gave, of course, the quantity of ice which had been converted into water; and from the weight of ice *melted* was calculated the specific heat of the body. Lavoisier and Laplace substituted a more complicated apparatus for the block of ice employed by Black, on account of the difficulty of obtaining ice in sufficiently large blocks. This apparatus is termed the ice calorimeter; the results obtained by it are inaccurate. Recently Bunsen has devised a calorimeter founded on the same principle, 'but in the use of which the sources of error are eliminated, and the physical constants determined with a degree of precision seldom before attained in researches of this kind.'

117. *Method of Mixtures.*—If a hot body be immersed in a cold liquid, after a short time both will be at the same temperature, the liquid gaining the same *quantity of heat* as the other body loses. This enables us to find the specific heat of a body in the following way: The body is weighed, and then raised to a known temperature, by keeping it, for instance, in a current of steam for some time; it is then immersed in *cold water*, the *weight* and *temperature* of which are known. Let:—

W=weight of the body.

w=weight of water.

H=specific heat of the body.

The specific heat of water is unity.

D=difference of the temperature of the body before and after the experiment.

d=difference of the temperatures of the water.

By par. 114, the quantity of heat lost by the body is WHD.

And the quantity of heat gained by the water is wd.

Therefore—

$$WHD=wd\text{ ; and }H=\frac{wd}{WD}$$

Example:—If 5 oz of gold at a temperature of 67° C. are immersed in 6 oz. of water at 29°, the temperature of the mixture becomes 30°; what is the specific heat of gold?

Here, $D=67^\circ-30^\circ=37^\circ$; and $d=30^\circ-29^\circ=1$

$W=5$; and $w=6$

We have—

$$\text{Sp. ht. of Gold}=H=\frac{w\,d}{W\,D}=\frac{6}{5\times 37}=0{\cdot}03243.$$

Exercises on Specific Heat.

27. Taking the specific heat of iron as 0·114, how much sensible heat would be absorbed by 10 lbs. of iron raised from 0° to 20° C.?

28. How much heat would be disengaged, if 15 lbs. of copper were cooled down from 100° to 5° C., its specific heat being 0·096?

29. If 3 lbs. of mercury, at 30° C., are mixed with 3 lbs. of water, at 10° C., and the temperature of the mixture is found to be 10·639° C., what is the specific heat of the mercury?

30. If 0·685 lb. of sulphur, at 60° C., is immersed in 4·573 lbs. of water, at 12° C. and the temperature of the mixture is 13·42°, what is the specific heat of the sulphur?

31. If 4 lbs. of zinc at a temperature of 97° C., be immersed in 11 lbs. of water at 5°, and the resulting temperature of the mixture be 8°, what is the specific heat of zinc?

32. If 13 lbs. of platinum at 50° be immersed in 6 lbs. of water at 36°, and the resulting temperature of the mixture be 37°, what is the specific heat of platinum?

118. It is obvious that the vessel which holds the cooling water must change its temperature with that of the water, so that the heat lost by the substance which is cooled not only raises the temperature of the water, but also the temperature of the material of which the vessel is composed to the same degree. In accurate experiments the amount of heat required for this purpose must be estimated and allowed for. The amount is easily estimated if we know the weight of the vessel and the specific heat of the substance of which it is made. The vessel is usually a small cylinder of silver or brass, with thin polished sides, so that these data can be readily obtained. If w' be the weight of the vessel and h' its specific heat, then the vessel, being raised through d degrees, will absorb a quantity of heat $w'\,h'\,d$. Now if v be the weight of water which would absorb as much heat as the vessel does, then $v=w'\,h'$, hence the heat lost by the body is absorbed by the water and the vessel, that is by a quantity of water w, and by an equivalent to a quantity of water v. Thus, instead of w in the preceding formula, we write $w+v$; the formula then becomes—

$$\mathrm{H}=\frac{(w+v)\,d}{\mathrm{W\,D}}\,;$$

v is said to be the reduced value in water of the vessel.

Example :—Three lbs. of silver at a temperature of 84° are immersed in 5¾ lbs. of water contained in a glass vessel whose weight is 1¼lb., both being at 10° ; the mixture becomes of a temperature of 12°. Calculate the specific heat of silver, that of glass being ·198?

Here $w'=1\tfrac{1}{4}=1{\cdot}25,\ h'={\cdot}198$; therefore, $v=w'\,h'=1{\cdot}25\times{\cdot}198={\cdot}2475$;

and $w=5\tfrac{3}{4}=5{\cdot}75,\ d=12-10=2$; $\mathrm{D}=84-12, \mathrm{W}=3$ hence :

$$\mathrm{H}=\frac{(v+w)\,d}{\mathrm{W\,D}}=\frac{({\cdot}2475+5{\cdot}75)\times 2}{3\times 72}=\frac{5{\cdot}9975\times 2}{3\times 72}={\cdot}055$$

119. In accurate experiments, it is necessary also to allow for the heat absorbed by the glass and mercury of the thermometer ; and if, as is usually the case, the substance is enclosed in a glass tube in a small basket of wirework, it is also necessary to pay regard to the weight and specific heat of these envelopes in the calculation. This method of determining the specific heat of solids and liquids admits of great accuracy, but its practical application requires many precautions, owing to the numerous sources of error.

Exercises on Specific Heat.

33. A piece of iron weighing 20 grammes at the temperature of 98° C. is dropped into a glass vessel weighing 12 grammes, and containing 150 grammes of water at 10° C. The temperature of the water is thus raised 10·29 C. Required the specific heat of the iron, knowing that the specific heat of the glass is 0·19768.

34. Calculate the specific heat of oil of turpentine from the following data :—42·57 grammes of the oil at 33·7 were mixed with 470·3 grammes of water at 12·23°. The temperature of the mixture was found to be 15·57°.

120. When we wish to determine the specific heat of a liquid or of a body which liquefies at a temperature little higher than that of the surrounding air, we are obliged to have recourse to the inverse method, which consists in cooling the body in a freezing mixture, and determining the depression of temperature which is produced by its immersion in the calorimeter.

121. The specific heat of a body is not a constant quantity; it increases with the temperature; a greater amount of heat is therefore required to raise a body one degree at a high, than at a low, temperature. Examples:—The specific heat of water at 180° C. is 1·0364; the mean specific heat of silver, between 0° and 100°, is 0·0557: while between 0° and 200° it is 0·0611. The student will now understand the reason why, in the definition of the unit of heat, a particular temperature, and not 1° at any temperature, was named.

122. The specific heat of a body is greater in the liquid than in the solid state: thus the specific heat of water is 1, that of ice is 0·504; liquid bromine, 0·111, solid bromine, 0·081; melted tin, ·0637, solid tin, ·05632. Knowing these facts, we might infer what is really the case,—that the increase of specific heat with the temperature is greatest as solid bodies approach the melting-point; and that the specific heat of solids diminishes with the increase in their density, or molecular aggregation. Examples:—Solid copper, when annealed, has a specific heat equal to ·09501; but by hammering, its density is increased, and consequently its specific heat is diminished. The specific heat of copper which has been hammered is ·0936. The density of tin and lead is not increased by mechanical pressure; their specific heats are not, therefore, diminished by pressure. The specific heat of a substance is not the same in its different allotropic modifications; in these cases, also, the specific heat diminishes as the density of the substance increases. Example:—

	Specific gravity	Specific heat
Wood charcoal .	. 0·300	0·2415
Graphite . .	. 2·300	0·2027
Diamond . .	. 3·500	0·1469

123. The specific heat of liquids, with the exception of water, increases with the temperature much more rapidly than that of solids.

124. Mercury, on account of its small specific heat, is rapidly heated and cooled. This property, as well as some of its other properties which have been already explained, renders it well adapted for a thermometric agent. Water, on the other hand, from its great specific heat, requires a long time in being heated or cooled. For the same weight and temperature, it absorbs or gives out far more heat than any other known liquid or solid, and is only exceeded by that of hydrogen. The ocean becomes, from this property of water, the great storehouse of heat; and we can at once see the important part water plays as the great regulator of climates, on account of this property. Comparing equal weights, the specific heat of water being 1, that of atmospheric air is 0·25; hence 1 lb. of water, in losing 1° of heat, would raise the temperature of 1 lb. of air 4°, or 4 lbs. of air 1°; but, owing to its much less density, 1 lb. of air fills about 770 times the space of 1 lb. of water; consequently, 4 lbs. of air fill 3,080 times the space; therefore, water, by parting with 1° of heat to the air, increases with that heat the temperature of 3,080 times its own bulk of air 1°. A large quantity of heat becomes stored up in the ocean in summer, which is slowly restored to the air in winter. 'Hence one cause of the absence of extremes in an island climate. In various parts of the Continent fruits grow which our summers cannot ripen; but in these same parts our evergreens are unknown; for they cannot live through the winters. The winter of Iceland is, as a general rule, milder than that of Lombardy.'

125. *Specific heat of gases.*—The specific heat of gases is referred either to that of water or air. Gases, according as they are allowed to *expand*, or they are *not* allowed to expand, being kept at a constant volume, on being heated, have a different capacity for heat; the former is called the *specific heat under constant pressure*, the latter *the specific heat under constant volume*; under the former condition the specific heat is greater than under the latter condition, by the amount of heat *consumed in the expansion.*

126. The specific heat of air, and of gases not readily liquefiable, does not increase with a rise of temperature, at all events between the temperatures of −30° and 200° C. It has also been ascertained that, for pressures ranging between one and twelve atmospheres, the specific heat of *equal weights* of air, and of gases not readily liquefiable, is uniformly the same, and is independent of the density; consequently, the specific heat of a given *volume* of gas increases directly as its density is increased.

127. All solid, liquid, and gaseous bodies, we have learned, *expand* when they are heated and *contract* when they are cooled; and the rate and degree of expansion of each substance are always the same, no matter how often it may be heated, if its temperature be the same at the commencement of each experiment, and the same amount of heat be added. For example,—if a substance at the temperature of 40° F. be heated up to 100° F., it will always expand to the same bulk when raised from 40° to 100° F., and it will always contract to the same bulk when cooled down again to 40° F.[1]

128. The same amount of heat affects matter in its three states very unequally. It produces the greatest expansion in gases, and the least in solids, because there is no cohesiveness between the particles of matter in the gaseous state; the heat has, therefore, no opposing force to overcome; and as the cohesive force is less in liquids than in solids, a less amount of heat is neutralised in overcoming this opposing force, and, therefore, for the same amount of heat, a greater expansion takes place in liquids than in solids.

129. All gases, on account of the absence of any cohesiveness between their particles, *expand alike*,[2] the pressure being the same, for equal additions of heat; and *contract alike* for equal subtractions of heat; but the different liquids and solids expand and contract unequally, for equal additions and subtractions of heat.

130. Having stated that all gases expand and contract alike on account of the absence of any cohesiveness between their particles, the following law, known as the law of Charles, that *all gases measured under constant pressure, expand equally for equal increments of heat*, might even be inferred; one volume of any gas at 0° C. expands to 1·3665 at 100° C., so that the coefficient of expansion of gases is 0·003665 or nearly $\frac{1}{273}$ for an increase from 0° to 1° C.

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[1] 'Lead is an exception to the rule; it is so soft that its particles slide over each other in the act of expansion, and do not return to their original position. For example, a leaden pipe used for conveying steam permanently lengthens some inches in a short time; and the leaden flooring of a sink which often receives hot water, becomes, in the course of time, thrown up into ridges and puckers.'—*Graham.*

[2] All that is stated with respect to gases applies also to vapours at some distance above their points of condensation.

131. We have learned that when substances are submitted to the action of heat, their temperature is *raised*, and they increase in *volume*. These are the first effects of heat on all substances, and they are the only ones produced on gases. But if the application of heat to solids and liquids be continued after these changes have commenced, a temperature is reached at which the solids and liquids change their state. The temperature at which each *solid changes its state is constant and peculiar to itself, and this is called its melting-point, or point of fusion*. Ebullition, like fusion, *for a given pressure commences at a certain temperature, which varies in different liquids, but which, for equal pressures, is always the same in the same liquid*; but this is not the case with evaporation, which takes place even with the same liquid at very different temperatures.

132. For the conversion of a solid into a liquid, and a liquid into a gas, a large amount of heat must be communicated to the solid, and to the liquid, to cause them to change their state; and this heat is *not* indicated in the substance by the thermometer; thus, if equal weights of water at 0° C., and water at 79° C., be mixed together, the temperature of the mixture will be the mean of the two temperatures, viz. 39·5°; but if 1 lb. of pounded ice at 0° is mixed with 1 lb. of water at 79°, the ice will be melted, but the temperature of the 2 lbs. of water will only be 0°; 79° of heat have, therefore, disappeared or become latent in melting the 1 lb. of frozen water. This absorbed heat is not lost or destroyed, for if the 2 lbs. of water be frozen that heat, and the latent heat of the water employed, will be evolved. The heat required to melt unit mass of ice without change of temperature is called the *heat of liquefaction of ice* or the *latent heat of water*, and amounts, as we have learned, to 79 thermal units of the Centigrade scale; an amount of heat that would raise 79 lbs. of water one degree Centigrade.

133. Black, in 1757, was the first to clearly establish that in order to convert water at the boiling temperature into steam of the same temperature a large quantity of heat must be communicated to it. 'He and many of his followers supposed heat to be a substance which when it makes a thing hot is sensible, but which when it is not perceived by the hand or the thermometer still exists in the body in a latent or concealed state. Black supposed that the difference between boiling water and steam is, that the steam contains a great deal more caloric than the hot water, so that it may be considered a compound of water and caloric; but, since this additional caloric produces no effect on the temperature, but lurks concealed in the steam ready to appear when it is condensed, he called this part of the heat latent heat.'

134. This view of the theory of heat, as has been previously explained, is not now accepted; and the late Professor Clerk Maxwell, in his work 'Theory of Heat,' which we quoted the last extract from, says, 'and, as I think Sir Isaac Newton's opinion that heat consists in the internal motion of the particles of bodies much the most probable, I choose to use the expression, "heat is generated," an expression Cavendish first employed.'

135. 'We shall not now be in danger of any error if we use latent heat as an expression meaning neither more nor less than this:

136. 'Definition.—*Latent heat is the quantity of heat which must be communicated to a body in a given state in order to convert it into another state without changing its temperature.*

137. 'We here recognise the fact that heat when applied to a body may act in two ways: by changing its state, or by raising its temperature; and that, in certain cases, it may act by changing the state without increasing the temperature.

138. 'The most important cases in which heat is thus employed are—

'(1) The conversion of solids into liquids. This is called melting or fusion. In the reverse process of freezing or solidification heat must be allowed to escape from the body to an equal amount.

'(2) The conversion of liquids (or solids) into the gaseous state. This is called evaporation, and its reverse condensation.

'(3) When a gas expands, in order to maintain the temperature constant, heat must be communicated to it, and this, when properly defined, may be called the latent heat of expansion.

'(4) There are many chemical changes during which heat is generated or disappears.

139. 'In all these cases the quantity of heat which enters or leaves the body may be measured, and in order to express the result of this measurement in a convenient form, we may call it the latent heat required for a given change in the substance.

140. 'We must carefully remember that all we know about heat is what occurs when it passes from one body to another, and that we must not assume that after heat has entered a substance it exists in the form of heat within that substance. That we have no right to make such an assumption will be abundantly shown by the demonstration that heat may be transformed into and may be produced from something which is not heat.'

141. The facility with which solids liquefy is proportional to the quantity of heat required for their liquefaction. Example: Ice cannot be liquefied until it has received as much heat as would raise its own weight of water 142·65° F.; therefore ice liquefies very slowly, because the *latent* heat of water is considerable; whilst phosphorus and lead are liquefied by as much heat as would raise their own weight of water 9° F.; therefore, they melt rapidly, because the heat required for their liquefaction is small.

142. Solids when heated up to their melting-point *invariably liquefy*; they cannot be heated beyond this point and remain solid. But liquids, under particular circumstances, may be cooled down *below* their freezing-point, or point of solidification, that is below the point at which, when in the solid state, the same substances melt. Example: Ice invariably melts or liquefies at 32° F. (0° C.), but water can, under certain circumstances, be cooled down to 5° F. and not freeze. This retardation of solidification can be effected by slow cooling, and protection from all agitation; but if it be the least agitated a portion of it is instantly converted into ice, and the temperature immediately rises to 32° F., by reason of the portion which becomes solid disengaging its latent heat, which becomes sensible, and keeps a portion of the water, at least for a time, in the liquid state. The heat disengaged would raise as much which performs this work is what Black called latent heat, because it is not sensible ; it only becomes sensible when the cohesive force, either by cooling or other agency, again comes into play.

149. The molecules of matter in each of the three states in which matter can exist are, it is supposed, in a continual state of motion ; in solids the movement of the molecules is confined to a very small region of space. In fluids the path of the molecules is not confined within a limited region, as in the case of solids, but may penetrate to any part of the space occupied by the fluid. In gases the molecules have very great freedom of motion, and they are supposed to move forward in straight lines till they impinge against each other, or against the sides of the vessel containing the gas; their course then is changed, and they start on a new path. These encounters also alter, and produce an inequality in, their velocity. 'This constant impact of the molecules produces the expansive tendency or elasticity which is peculiar to the gaseous state. The total quantity of heat in the gas is made up of the progressive motion of the molecules, together with the vibratory and other motions of the constituent atoms ; but the progressive motion alone, which is the cause of the expansive tendency, determines the temperature. Now, the outward pressure exerted by the gas against the containing envelope arises, according to our hypothesis, from the impact to a great number of gaseous molecules against the sides of the vessel. But at any given temperature—that is, with any given velocity—the number of such impacts taking place in a given time, must vary inversely as the volume of the given quantity of gas ; hence *the pressure varies inversely as the volume or directly as the density*, which is Boyle's law.

150. 'When the volume of a gas is constant, the pressure resulting from the impact of the molecules is proportional to the sum of the masses of all the molecules multiplied into the squares of their velocities—in other words, to the so-called *vis viva* or *living force* of the progressive motion.

151. 'If, for example, the velocity be doubled, each molecule will strike the sides of the vessel with a twofold force, and its number of impacts in a given time will also be doubled ; hence the total pressure will be quadrupled.

152. 'Now we know when a given quantity of any gas is maintained at a constant volume, it tends to expand by $\frac{1}{273}$ of its bulk at zero for each degree Centigrade. Hence the pressure or elastic force increases proportionally to the temperature reckoned from $-273°$ C.; that is to say, to the absolute temperature. Consequently, the *absolute temperature is proportional to the* vis viva *of the progressive motion.*' We may notice in conclusion that the expansion of air is attended with a diminution of temperature; and its compression with an increase of temperature; in fact, a compression or expansion amounting to 0·00366 of the volume which the air would occupy at zero, alters its temperature by 0·41° C. The cooling of the air by expansion, and the development of heat by its compression, is thus explained on the dynamical theory of heat. Work is done in the expanding; consequently a portion of the heat of temperature is expended in producing the energy of expansion; hence the diminution of the temperature of the expanded gas. On the other hand, on compression there is a diminution of the expansive force, and to this diminution is due the development of heat. At times a cloud is formed in the receiver of the air-pump, when the air is being pumped out of it, because the rarefied air becomes chilled; consequently some of the aqueous vapour it contained is condensed and forms the cloud; if the air, during its presence, be readmitted, as the air necessarily becomes heated, 'the cloud melts away, and the perfect transparency of the air is restored.'

Examination Questions.

55. Define what is meant by the mechanical equivalent of heat; state what it is in foot-pounds.

56. In exhausting the receiver of an air-pump a cloud sometimes appears in the receiver; explain the cause of this.

57. Explain the difference between the specific heat of air at constant pressure, and its specific heat at constant volume. Which of them is greater, and why?

58. What is meant by saying that the specific heat of water is 30 times as great as that of mercury?

CHAPTER III.

WEIGHT—THE BALANCE—WEIGHTS AND MEASURES—EXERCISES—WEIGHT AND PRESSURE OF THE AIR—DENSITY OF THE AIR AT DIFFERENT HEIGHTS—THE AIR-PUMP—EXPERIMENTS—FILTER-PUMP—PHYSICAL PROPERTIES OF GASES—EXPERIMENTS—EXERCISES—QUESTIONS.

153. THE particles of different bodies are, as we have already stated, attracted to each other, and this attraction is *all-pervading*, for it operates throughout space, and is termed *gravitation*. Now as each particle attracts, the *mass* containing the greatest number of particles possesses, in consequence, the greatest attractive force; thus, according to the law of Newton, *the attractive force of bodies is to each other directly as their masses*, and *inversely as the square of their distances*; therefore, if the mass of one body be two, three, four, or five times *greater* than that of another, its gravity or attractive force is also two, three, four, or five times *greater*; and with a *distance* of two, three, four, or five times greater, the force of gravity will be four, nine, sixteen, or twenty-five times *less*.

154. Owing to this attractive force, the earth attracts all terrestrial objects towards itself, and is itself attracted by them; but as their masses are so extremely small in comparison with the earth's mass, their motion only is apparent to us; hence the earth's attraction, or terrestrial gravitation as it is generally termed, tends to cause all bodies to fall to its surface. One of the effects of this attraction is the downward pressure which substances exert on whatever supports or prevents them from falling to the earth, and *this pressure is termed their weight*, or more precisely their absolute weight; and this is one of the subjects we have to consider in the present chapter.

155. Mass, or the quantity of matter in a substance, is estimated in terms of some standard: the British unit of mass is the standard pound (avoirdupois); and the balance, or weighing-machine, is used for the comparison of masses. There are other kinds or relations of weight; these we shall treat of by-and-by.

156. The determining of the weight of substances is being constantly carried on in every-day life, and 'scales' is the name usually given to the smaller weighing-machines in common use.

157. The weighing-machine used in experimental science is called a balance (fig. 25); it has to be not only extremely accurate, as also the weights employed, but must also be so delicate that it will readily indicate 0·002 of a grain when carrying a load of 1,000 grains; for, as the late Dr. Miller remarked, the foundation of all accuracy in experimental science consists in the possibility of determining with exactness the quantity or mass, and the size or bulk, of those substances which are submitted to examination. The late Baron Liebig made the following observations in his 'Letters on Chemistry' in reference to this instrument. 'The great distinction between the manner of proceeding in chemistry and natural philosophy is, that one *weighs*, the other *measures*. The natural philosopher has applied his measures to nature for many centuries, but only for fifty years have we attempted to advance our philosophy by *weighing*.

158. 'For the great discoveries of Lavoisier, he was indebted to the "balance," that incomparable instrument which gives permanence to every observation, dispels all ambiguity, establishes truth, detects error, or shows us that we are in the true path.

159. 'The balance, once adopted as a means of investigating nature, put an end to the reign of Aristotle in physics. The explanation of natural phenomena by mere fanciful speculations, gave place to a true natural philosophy.'

160. It would be out of place in a work like this to enter upon the mechanism of the balance, but it will not be out of place to give the student a few hints as to the using of a delicate chemical balance. It is always placed in a glass case (fig. 25), in order to protect it from air-currents, dust, and moisture; as air-currents affect the weighing, no substance whilst hot must be weighed, because heat causes air-currents. An index-needle or pointer is fixed to the beam, and oscillates with it in front of the graduated arc *a*, and when the beam is suspended and at rest, no weight of any kind being in the scale-pans, the index should point to the zero of the scale. When the correct weight is being nearly approached, the balance-case should be closed; and as the beam of a delicate balance oscillates for a long time, it is usual to weigh, in order to save time, whilst it is oscillating, adjusting the weights until the movements of the pointer on each side of the zero of the scale are equal; results even more accurate, are obtained in this way, than if the beam was at rest during the time of weighing, because whilst oscillating, we are certain that no resistance is offered through the beam 'sticking.' The same pan for the weights should always be used; the two arms of the beam should be of equal length, otherwise unequal weights would be required to produce equilibrium. The error occasioned by inequality of the arms of the balance may be eliminated by the method of *double* weighing, which consists in weighing the substance first in one pan and then in the other; the weights will not be the same in the two weighings, the geometric mean is, therefore, the correct weight, and is found by multiplying the true apparent weights together, and taking their square root; but if the true apparent weights be nearly equal, their common arithmetic mean is quite close enough to the truth. 'The balance will only give accurate comparisons of mass when the weighing is performed *in vacuo*, unless the body weighed happens to have the same density as the standard weights. The reason is that if the body and the weights are of unequal volume, they will displace different amounts of air, and hence lose weight unequally. It is, however, easy to make the requisite correction.' When the weighing is completed, the weights must always be removed from the scale-pan and replaced in the weight-box.

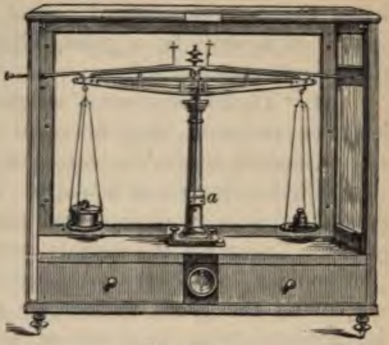

Fig. 25.

161. The standard of weight used in this country is an imperial standard called the avoirdupois pound, which is subdivided into 7,000 grains. It was enacted in 1855, "that the platinum weight deposited in the Exchequer shall be denominated the imperial standard pound *avoirdupois*, and that the 7000th part shall be a grain, while 5,760 such grains shall make the pound troy."

162. The two systems of weights—troy and avoirdupois—have no common integer except the *grain*. Although the terms, pound, ounce, are common to both systems, they denote different quantities in each. The English *troy* pound is subdivided into 12 ounces, and each ounce is equal to 480 grains. These are the weights generally employed: but for scientific purposes ambiguity is most easily avoided by employing the *grain* as integer, and the laboratory should be provided with good sets of weights, from 1,000 grains downwards. The grain should be subdivided into tenths, hundredths, and thousandths.

TROY WEIGHT.

Pound.	Ounce.	Pennyweight.	Grain.
		dwts.	gr.
1	= 12	= 240	= 5760
	1	= 20	= 480
		1	= 24

AVOIRDUPOIS WEIGHT.

Pound.	Ounces.	Drachms.	Grains.
1	= 16	= 256	= 7000
	1	= 16	= 437·5
		1	= 27·34

163. The system of weights is connected with the measures of capacity in use in this country, through the medium of the imperial gallon, which is defined by an Act of Parliament of the year 1824 to be a measure containing 10 lbs. avoirdupois of distilled water weighed in air at a temperature of 62° F., the barometer standing at 30 inches. The gallon of distilled water, therefore, contains 70,000 grains.

164. These measures of capacity are related to those of length by the determination that a gallon contains 277·276 cubic inches. A cubic inch of distilled water weighs, in air at 62° F., with the barometer at 30 inches, 252·456 grains; *in*

vacuo it weighs 252·722 grains. The standard of length is the yard measure, and is subdivided into 36 inches. The standard yard is defined by an Act of Parliament passed in 1855, whereby it is enacted that the straight line or distance between the centres of two gold plugs in the bronze bar deposited in the office of the Exchequer shall be the genuine standard yard at 62° F., and if lost it shall be replaced by means of its copies.' —*Miller.*

165. The French or metrical system of weights and measures is connected together in a more philosophical manner than the English one; it is the one that is almost universally adopted abroad by scientific men, and is now largely employed by scientific men in England. It is a decimal system, founded upon the mètre, which is $\frac{1}{10,000,000}$th part of a quadrant of the meridian: this constitutes the *unit* of the system.

166. The French mètre is equal to 39·3707 English inches. The *mètre* is, in France, the integer of the measure of *length*, and from it all measures of surface, capacity, and weight are derived. The integer of the measure of capacity is the *litre*, which is the cubic decimètre, and is equal to 35·275 fluid ounces, or 1·763 imperial pints. The integer of the measure of weight is the *gramme*,=15·43 English grains. It is exactly equal to the weight of a *cubic centimètre* of water, weighed *in vacuo* at its maximum density (39·2° F.). The cubic centimètre is employed by French chemists in all measurements of gases, in place of our cubic inch. It is equal to 0·06103 of a cubic inch. The weight of the cubic centimètre of water is to the cubic inch of water as 15·432 to 252·468; hence there are 16·34 cubic centimètres to an English cubic inch.

167. This rule is sufficiently correct for practical purposes. It is to be observed, however, that the French take the weight of the cubic centimètre of water at 39·2° F. *in vacuo.* The English take the weight of the cubic inch (252·458) grains at 62° F. in air. Assuming the specific gravity of water at 32° to be 1·000,000, the specific gravity at 39·2° is to the specific gravity at 62° as 1·000,099 to 0·999,000.

168. 'The French measures increase and decrease in decimal proportions. For the *increase*, a prefix is used derived from the Greek *deca*, *hecto*, *kilo*, and *myria*; the integer, whether mètre, litre, or gramme, being multiplied by 10, 100, 1,000, and 10,000 respectively. To indicate the *decrease*, the prefixes *deci*, *centi*, and *milli*, derived from the Latin, are employed. In this case the integer is supposed to be divided by 10, 100, or 1,000.

169. 'Various plans have been devised for converting the French weights and measures into their English equivalents. The following tables will be found useful for this purpose :—

Measures of Length.

	In English inches		In English inches
Millimètre . =	·03937	Decamètre =	393·70790
Centimètre . =	·39371	Hectomètre =	3937·07900
Decimètre . =	3·93708	Kilomètre . =	39370·79000
Mètre . . =	39·37079	Myriomètre =	393707·90000

1 inch = 2·539954 centimètres.
1 foot = 3·0479449 decimètres.

1 yard = 0·9143835 mètre.
1 mile = 1·6093149 kilomètre.

Measures of Volume, or Capacity in Cubic Inches.

Millilitre, or cubic centimètre . . . . . =	0·06103
Centilitre, or 10 cubic centimètres . . . . =	0·61027
Decilitre, or 100 cubic centimètres . . . . =	6·10271
Litre, or cubic decimètre, or 1000 cubic centimètres . . . . =	61·02705
Decalitre, or centistere . . . . . . =	610·27052
Hectolitre, or decistere . . . . . . =	6102·70515
Kilolitre, or stere, or cubic mètre. . . . =	61027·05152
Myriolitre, or decastere . . . . . . =	610270·51519

1 cubic inch . . = 16·386176 cubic centimètres.
1 cubic foot . . = 28·315312 cubic decimètres, or litres.
1 gallon . . . = 4·543358 litres.

Measures of Weight.

	In English grains		In English grains
Milligramme . =	0·01543	Decagramme =	154·32349
Centigramme . =	0·15432	Hectogramme =	1543·23488
Decigramme . =	1·54323	Kilogramme =	15432·34880
Gramme . . =	15·43235	Myriogramme =	154323·48800

1 grain = 0·064799 gram.
1 troy oz. = 31·103496 grams.

1 lb. avois. = 0·453593 kilogr.
1 cwt. . = 50·802377 kilogrs.

MEASURES OF SURFACE.

	In English square feet	In English square yards = 9 square feet	In English poles = 272·25 square feet	In English roods = 10,890 square feet	In English acres = 43,560 square feet
Centiare or square mètre . .	10·764299	1·196033	0·0395383	0·0009885	0·0002471
Are or 100 square mètres . .	1076·429934	119·603326	3·9538290	0·0988457	0·0247114
Hectare or 10,000 square mètres .	107642·993418	11960·332602	395·3828959	9·8845724	2·4711431

1 square inch = 6·4513669 square centimètres.
1 square foot = 9·2899683 square decimètres.
1 square yard = 0·83609715 square mètre, or centiare.
1 acre = 0·40467102 hectare.

170. 'With respect to the equivalents of English and French weights and measures, some slight differences will be found among English writers. These arise from the calculations being based on the employment of a larger or smaller number of decimal figures.'

Exercises on the Conversion of French into English Measures, and vice versâ.

59. Convert five centimètres into the corresponding length in English inches.

60. What is the corresponding French weight to one English grain?

61. Convert twelve cubic centimètres into English inches.

62. What is the corresponding French volume to three cubic inches?

63. Convert three kilogrammes into the corresponding English avoirdupois weight.

64. Convert one English inch into the corresponding French measure.

65. I determine the amount of silver in a silver-lead ore, I obtain 0·020 grains of silver from 400 grains of the ore; for commercial purposes the silver contained in any given ore is estimated in ounces, pennyweights, and grains; one ton of the ore being, in all cases, taken as the standard of unity. How many ounces, pennyweights, and grains of silver were contained in a ton of this ore?

66. The late Baron Liebig, in his work *The Natural Laws of Husbandry*, states that an average wheat crop yields from a hectare of land 2,000 kilogrammes of grain and 5,000 kilogrammes of straw; and a rye crop from the same area 1,600 kilogrammes of grain and 3,800 kilogrammes of straw; convert these foreign weights and measures into their English equivalents.

171. All bodies in each of the three states of aggregation, possess weight, and *dissimilar* bodies have almost invariably a different weight for equal volumes. A child knows that solids and liquids possess weight, but it is not so readily realised that gases possess weight, and that they *differ* in weight for equal volumes; in other words, all gases have not the same weight; some are heavier, some are lighter than others; in brief, gases differ in weight like solids and liquids. The ascension of a balloon in the air is a very instructive illustration that gaseous substances differ in weight; for its ascension is due to its being filled with a gas lighter than atmospheric air.

Experiments on the Weight of Gases.

67. Fill a small balloon of goldbeater's skin, or a soap-bubble,[1] with dry hydrogen gas, or coal-gas; after the balloon is filled allow it to ascend in the air; hydrogen being much lighter in weight than coal-gas, the balloon will be much more buoyant if filled with it, than if filled with coal-gas.

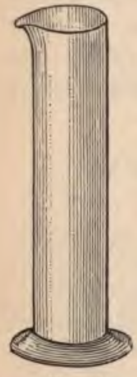

FIG. 26.

68. Prepare some carbon dioxide and let the exit tube of the apparatus dip into a dry glass vessel (fig. 26); the tube ought to reach to the bottom of the vessel. As soon as the vessel is filled with the gas, which is ascertained by a candle being extinguished when introduced just below the edge of it, pour the gas as you would a liquid into another vessel of exactly equal size and shape. Prove that the gas has been really transferred from the one vessel to the other by introducing a lighted candle first into the vessel originally containing it (in which the candle ought now to burn), and then into the one which ought now, if the experiment has been properly performed, to contain it, and in which the candle ought not now to burn. If great care be observed in the experiment, it may be transferred from one vessel to the other three or four times. This experiment shows the great density of the gas: it also shows the incapability of the gas to support the combustion of a candle.

69. Fill a wide and deep glass jar half-full of carbon dioxide, the upper half being filled with atmospheric air; fill a collodion balloon with hydrogen and air in such proportions that the balloon will just sink in air; when so prepared, let the balloon sink into the jar; as soon as it reaches the uppermost layer of the oxide of carbon it will rebound as if it had touched a solid body; it will finally float quietly upon it, and it is curious to see the balloon suspended midway in the jar, apparently resting upon

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[1] For making soap-bubbles the following is the best method for making the soap solution: put into a quart bottle some of the best white Castile soap (or, still better, pure palm-oil soap). Cut the soap (about four ounces) into thin shavings, then fill up the bottle with distilled water or rain water, shake the mixture very well; repeat the shaking until a saturated solution of soap is obtained. If on standing the soap is not all dissolved, pour off the liquid, and add more water to the soap which remains undissolved, shaking as before. Then add to two volumes of the soap solution one volume of pure, concentrated glycerine.

nothing. This experiment is an exceedingly interesting and instructive one if well performed.

172. The barometer is an instrument that proves that the air possesses weight, for it is the pressure of the atmosphere on the surface of the mercury in the cistern (fig. 27) that sustains or makes the mercury in the tube stand about 30 inches above the level of that in the cistern, for if, instead of having the upper end of the tube hermetically sealed, it were closed by means of a stopcock, and that it was opened when the mercury was standing in the tube, the mercury would instantly sink in the tube to the *same level* as that in the cistern; proving that it was the pressure of the air upon the mercury in the cistern that supported the mercury column.

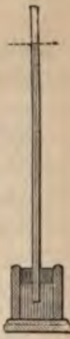
FIG. 27.

173. The barometer illustrated in fig. 27 is called the cistern barometer; there are other forms of the mercurial barometer, but they all depend on the same principle, viz. the pressure of the air on the mercury. There are barometers, as the aneroid, without any liquid; but they likewise depend on the pressure of the air.

174. The space above the column of mercury must be entirely free from *air* and *moisture*, for if present, they would exercise a pressure on the mercury in the tube which would counteract, to a greater or less extent according to their amount, the pressure of the air on the mercury in the cistern, and thus render the barometer more or less inaccurate. It is therefore necessary in the construction of a barometer, in order to have it accurate, to have not only the mercury pure and free from air, but the tube must also be free from air and moisture; to obtain this result, a small quantity of pure mercury is placed in the tube and boiled for some time. It is then allowed to cool, and a further quantity previously warmed, added, which is boiled, and so on, until the tube is quite full; in this manner the moisture and the air which adhere to the sides of the tube pass off with the mercurial vapour.

175. That portion of the upper part of the tube which is unfilled with mercury is the nearest approach to a perfect vacuum which can be procured by art, for on lowering the tube deep enough into the cistern of mercury, the whole tube will become completely filled with the mercury; and on elevating the tube in the cistern the mercury again sinks in the tube. The vacuum thus obtained is termed the Torricellian vacuum, from its being first observed in 1663 by Torricelli, a pupil of Galileo. Although this is the most perfect vacuum that can be obtained, it is, nevertheless, not a perfect one, as the apparently empty space of the tube contains a *minute* quantity of mercurial vapour.

176. It need scarcely be remarked that all other liquids would, like mercury, be supported in a tube by the pressure of the air, the height of the column varying with the weight of the liquid; water, for example, being 13½ times lighter than mercury, the height of a column of that liquid would be 13½ times greater than that of mercury: the column of water would therefore measure about 34 feet; in other words, when any other liquid than mercury is used, the height of the barometric column will be *inversely* as the *specific gravity*; the specific gravity of mercury at 60 F. is 13·568, the specific gravity of water being taken at 1·000; the height of a column of water corresponding to a column of 30 inches of mercury at 60° F. would therefore stand—

$$30 \times \frac{13{\cdot}568}{1{\cdot}000}, \text{ or } 407{\cdot}04 \text{ inches or } 33{\cdot}92 \text{ feet high.}$$

177. From these facts we learn that the pressure of the atmosphere is capable of supporting a column of mercury about 30 inches in height, and a column of water about 34 feet in height; or, to state it conversely, the pressure of a column of mercury of 30 inches, and a column of water of 34 feet, in height is equal to the pressure of the atmosphere.

178. As the pressure of the atmosphere supports a column of mercury whose base is a square inch 30 inches in height; and as a cubic inch of mercury weighs, in round numbers, half a pound, 30 cubic inches of the metal will weigh 15 lbs.; it, therefore, follows that the pressure of the atmosphere on each square inch of surface is equal to 15 lbs., and this pressure of 15 lbs. on the square inch is called the pressure *of one atmosphere.* If the elastic force of a gas or vapour, say that of steam in a boiler, were equal to a pressure of 60 lbs. on the square inch, then its pressure would be equal to that of four atmospheres, and it would be so stated.

179. Mercury being the densest of all liquids, it stands at the least height; it is therefore chosen in preference to any other liquid in constructing barometers.

180. Having stated in general terms how the barometer is constructed, and the principle in its construction for measuring the pressure or weight of the atmosphere, we must now describe it a little more fully, as it is an important instrument in the chemical laboratory, for at the moment the chemist weighs or measures a gas he must take the pressure of the atmosphere by means of this instrument; and he likewise requires to know the atmospheric pressure when he has to make very accurate weighings of light bodies, in order that the weight of air they displace may be allowed for.

181. As the height of the mercurial column varies, so does the quantity of mercury in the cistern vary; for example, if the column stands high, there must necessarily be less mercury in the cistern than when the column stands comparatively low; consequently, the true variation in the length of the column is not obtained by noticing the top of the mercury only, since the base of the column also varies, and therefore a correction has to be made for the amount of variation; but this correction is not needed in most of the best forms of the instrument now made for scientific purposes; as the whole scale is made movable, and, therefore, allows of adjustment as the level in the cistern *fluctuates*, consequently, the numbers marked on the scale always express *the real distance from the level of the mercury in the cistern.*

182. There are other corrections that require to be made which we must notice: the first we will allude to arises from the capillary action between the glass and mercury; this will be noticed in the chapter (VII.) which treats of adhesion; the next correction to be noticed is that for temperature. Mercury expands with increase of temperature, and contracts with decrease of temperature, hence its density changes, and inversely with its expansion, and therefore the height of the column will be inversely proportional to the specific gravity. To ascertain the temperature of a barometer a thermometer is always attached to it. All readings have, therefore, to be reduced to a determinate temperature; the temperature which has been adopted is that of melting ice, at which temperature the specific gravity of mercury is 13·6. Tables have been drawn up for making the necessary correction at different temperatures.

183. Many more examples might be given to prove that the air possesses weight; but we will notice first the instruments already illustrated, and which depend for their action on the pressure of the air. The action of the pipette (Exper. 5) depends upon atmospheric pressure; the lower end of the instrument is dipped into a liquid, and suction is applied by the mouth or other means to the upper end; the withdrawing of the air diminishes the pressure within; the liquid is therefore pressed upwards by the excess of the external over the internal pressure; it rises in the pipette, and on closing air-tight the upper opening with the finger the liquid cannot flow back even when the instrument is raised out of the liquid, provided the lower aperture is not too large; but if it is so large that the air and liquid have sufficient space to pass each other, then the latter cannot, of course, be retained in the pipette.

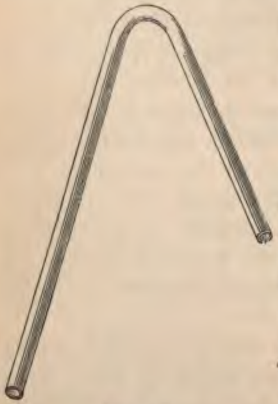

FIG. 28.

184. The syphon is another instrument of frequent use in the chemical laboratory. It is a bent glass tube open at both ends, and is generally shaped somewhat in the form of a U, and it is usually so bent that the legs are of unequal length. It is used for transferring a liquid from one vessel to another. It may be filled by inverting it, and so inclining it that the orifice of the shorter leg is higher than the other; the liquid is poured in until the longer leg is filled; its orifice is then closed with the finger, and more liquid is then added until the other leg is full; having closed both orifices with the fingers insert the shorter leg into the liquid which has to be removed, then withdraw the fingers and the liquid will flow out of the longer leg. It is frequently made to work by inserting one leg in the liquid and sucking at the other end. It is by the combined action of the *pressure* of the air and the *weight* of the liquid that the instrument acts. The pressure of the column of liquid in the immersed portion is exactly counterbalanced by the pressure of the water in the vessel; and the pressure of the air is equally exerted on the liquid in the two legs; in the leg immersed the pressure is transmitted by the liquid in which it is inserted; in the other the pressure acts directly on the aperture. For the syphon to act, the *length* of the column of liquid in the leg immersed from the surface of the liquid in the vessel to the highest point in the *bend* of the syphon must be less than the length of the column in the other leg; otherwise, the pressure of the air being equal on both columns of liquid, and these being of equal length, there could be no flow from the outer limb. If therefore the two legs are of equal length, one leg must be immersed in such a manner that the column of the liquid in the two limbs is of unequal length.

185. There is, of course, a limit to the action of the syphon, for in order to convey the liquid the bend of the syphon must not be at a greater height above the level of the liquid than that of a column of the liquid which the pressure of the atmosphere will support. We have learned already that it will support a column of water 34 feet in height, and a column of mercury 30 inches in height; therefore, the shorter limb should be rather less in length, if water has to be conveyed, than 34 feet from the level of the water to the bend, and rather less than 30 inches if mercury has to be conveyed; for the *height* to which a liquid can be raised by a syphon *depends on its specific gravity*.

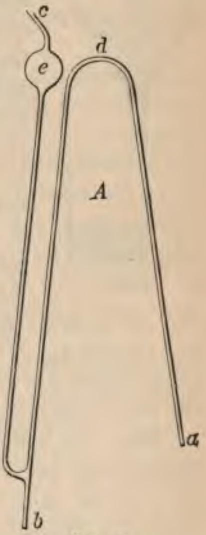

FIG. 29.

186. Another form of the syphon is represented in fig. 29. It is filled by dipping the end *a* into the liquid, and closing the end *b* with the finger. The syphon is then filled by sucking at *c*. When both branches, *ad* and *bd*, are full, the finger is withdrawn, and the liquid runs out at *b*. When syphoning acids the end *b* ought to be provided with a stopcock, so that the finger need not be employed.

187. The gas jar filled with water or mercury, and standing on the shelf of the pneumatic trough, (par. 54, fig. 8) is another example of the pressure of the air ; so also is the gasometer (fig. 9) when it is filled with water, and the stopcocks C and D are closed, and the nut of the tube F is opened ; no water flows from F, owing to the pressure of the air.

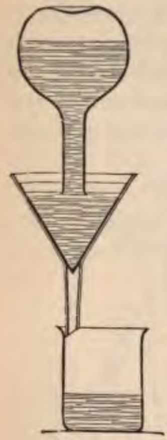

FIG. 30.

188. A simple contrivance sometimes adopted by chemists for washing solid substances on filters without having to give frequent attention to the washing, is founded on the same principle. A bottle is filled with water, and its mouth being closed air-tight by the finger is inverted into the filter so that its mouth dips about an inch into it, as shown in fig. 30. As long as the mouth remains covered with the liquid no water escapes from it, but whenever so much of the liquid has passed through the filter as to make the surface sink lower than the mouth of the bottle, the air ascends in the latter, and the water flows out until the mouth again becomes covered with the liquid, and this intermittent flow of water out of the bottle goes on until all the water has flowed out of it.

189. The action of the boy's toy called a sucker is due to atmospheric pressure. And flies and some other animals are able by the same action as that of the sucker to sustain their bodies in opposition to the force of gravity.

190. It is thought that all throughout infinite space there is no such thing as a vacuum ; all space is believed to be everywhere filled with matter of one form of aggregation or another ; in other words there is no such thing as an *empty space* in nature.

191. On the earth and within the limits of the earth's atmosphere, all space, that is not filled with matter in the solid or liquid state, is filled with atmospheric air, with a few trifling exceptions, as, for instance, the accumulations of one of the constituents—carbon dioxide, of the air in some valleys, and the accumulation of gases, other than air, in some chemical manufacturing processes, as, for example, the accumulation of carbon dioxide in the fermenting vessels in breweries ; but in the general sense the statement is true, that on the earth all space, that is not filled with solid or liquid matter, is filled with air. We have, therefore, next to consider some of the physical properties of the air.

192. The air is a mixture of several gases, viz., oxygen, nitrogen, carbon dioxide, ammonia, and a form of oxygen called ozone ; and it always contains a varying proportion of aqueous vapour : there is also present in it finely divided solid particles of different substances which it is not necessary in this part of the work to specially name.

193. The atmosphere has, it is believed, a *definite limit* ; from observations made as to the length of time during which twilight is visible in the zenith, it is considered that it reaches in a state of *sensible* density to a height of from forty to forty-five miles above the earth's surface. 'Meteorites, however, ignite at an elevation of about 200 miles, proving the presence of a medium, which, though of too great tenuity to reflect light, still possesses density, and offers resistance to the passage of bodies through it. It is probable that even this height does not denote the upper limit of the atmosphere.'

194. The density of the air is not the same throughout its entire height, for, owing to its elasticity, its density diminishes as its distance from the earth increases, the lower strata having a much greater density than the higher strata, because each layer bears the pressure of the layers above it, but not of those underneath ; therefore, the pressure must decrease on each ascending layer ; the lowest layer will bear the pressure of all the rest ; the one next above the lowest will bear the pressure of all but the layer underneath it ; and the third lowest layer will bear the pressure of all but the two layers underneath it ; and so on with the succeeding higher layers.

195. As the pressure on each ascending layer decreases so the density must decrease, because *the density of a gas is directly as the pressure, and inversely as the volume.* The air decreases in density in the following progression: at an elevation of 3·4 miles it is $\frac{1}{2}$ the density of the air at the sea-level; at 6·8 miles it is $\frac{1}{4}$; at 10·2 miles $\frac{1}{8}$; at 13·6 miles $\frac{1}{16}$; and at 17 miles $\frac{1}{32}$ of that density. The *bulk* of equal weights of air at these successive heights *increases* in geometrical progression, the volume being doubled for each step in the ascent; while the *density*, and the corresponding height of the barometer, *decrease* in the same geometric ratio, being at each successive elevation exactly half what they were in the preceding one; for all practical purposes it is sufficient to take it, that the barometer falls *one-tenth* of an inch for every 95 feet of height in this country in summer, and for heights not exceeding 3,000 feet. The barometer can be, and is, employed to measure the heights of mountains, and other altitudes.

196. Not only does the pressure of the atmosphere decrease as we ascend from the level of the sea, but at the sea-level the mean pressure of the atmosphere varies with the latitude. At the equator, the mean level of the barometer is 29·84 inches; it increases with the latitude up to 40° N. where it amounts to 30·04 inches; and at higher latitudes it again decreases.

197. But in addition to these changes of atmospheric pressure with change of altitude or place, the pressure at the same place undergoes continual variations. 'It is observed that the course of the barometer is generally in the opposite direction to that of the thermometer; that is, that when the temperature rises the barometer falls, and *vice versâ*; which indicates that the barometric variations at any given place are produced by the expansion or contraction of the air, and therefore by its change in density. If the temperature were the same throughout the whole extent of the atmosphere, no currents would be produced, and, at the same height, the atmospheric pressure would be everywhere the same. But when any portion of the atmosphere becomes warmer than the neighbouring parts, its specific gravity is diminished, and it rises and passes away through the upper regions of the atmosphere, whence it follows that the pressure is diminished, and the barometer falls. If any portion of the atmosphere retains its temperature, while the neighbouring parts become cooler, the same effect is produced; for in this case, too, the density of the first-mentioned portion is less than that of the others. Hence, also, it usually happens that an extraordinary fall of the barometer at one place is counterbalanced by an extraordinary rise at another place. With reference to the daily variations, they appear to result from the expansions and contractions which are periodically produced in the atmosphere by the heat of the sun during the rotation of the earth.'—*Ganot.*

198. Before commencing to describe the physical properties of air, *taking air as the type of gaseous bodies in general*, we must describe how air can be removed from a confined space, a vacuum being produced; this is effected by means of an instrument called the air-pump, which was invented by Otto von Guerick in 1650, a few years after the invention of the barometer.

199. There are various forms of this instrument, but the principle is the same in every form of it. It is identical in principle with the water pump; and the principle of its construction can be very easily explained and illustrated:—A metallic cylinder (fig. 31), accurately bored, is fitted with a piston which is covered with a flap or valve of oiled silk which opens upwards; when the piston is forced down, the air in the outer cylinder will, owing to the compression, escape into the atmosphere through the valve; on raising the piston no air will be able to re-enter the cylinder, and owing to the pressure of the atmosphere on its upper surface, a resistance will be experienced in raising it. If the cylinder has a valve at the bottom, opening in the same direction as that in the piston, as shown in the figure, this valve will on thrusting down the piston be closed by the elasticity of the air in the cylinder, while the upper valve will be opened; on raising the piston the effect is reversed—the valve in the bottom of the cylinder will be opened by the pressure of the air under it, and the air will, of course, pass through it.

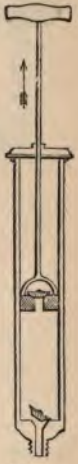

Fig. 31.

200. The air-pump, of which a perspective view is given in fig. 32, and a section in fig. 33, is an apparatus consisting usually of two exhausting syringes, B B′ (fig. 33), mounted so as to be worked by a single winch and handle, as represented at D, and communicating by a common pipe, T, with a glass vessel, R, in which may be placed the objects of experiment. The vessel R, called a receiver, has an edge, S, ground smooth, resting upon a plate, also ground smooth, and kept in air-tight connection with it by being smeared with lard. A stopcock C is provided in the pipe T, by which the communications between the receiver R and the syringes can be made and broken at pleasure. Another stopcock is provided elsewhere, by which a communication can be made at pleasure between the interior of the receiver R and the external air. To indicate the extent to which the rarefaction is carried from time to time by the operation of the syringes, a mercurial gauge, H G M, is provided, constructed in all respects similar to a barometer. The atmo-

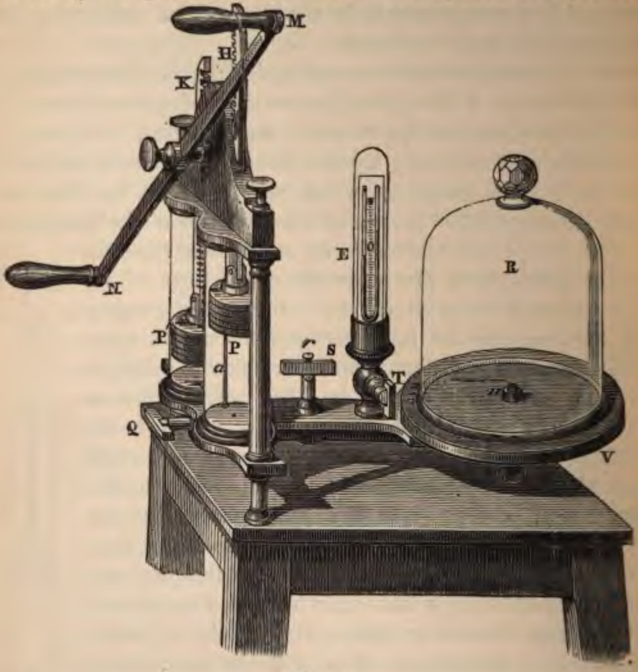

FIG. 32.

sphere presses on the surface of the mercury in the cistern M, while the column of mercury in the tube HG is pressed upon by the rarefied air in R. The height of the column, therefore, sustained in the tube, indicates the difference between the pressure of the external air and the air in the receiver.

'When a gauge of the form represented in fig. 33 is used, it is necessary that it should have the height of about 30 inches, since when a high degree of rarefaction has been effected, a column of mercury will be sustained in the tube HG very little less than in the common barometer. In small pumps, where this height would be inconvenient, a syphon-gauge, such as that represented in fig. 34 is used. This gauge is screwed on to a pipe communicating with the receiver. Mercury fills the leg A B, which is closed at the top A, and partially fills the leg S. When the atmosphere communicates freely with the tube D C, the surface of the mercury in S, being pressed by its full force, sustains all the mercury which the tube B A can contain, and this tube consequently remains completely filled; but when the pipe D C S is put in communication with the exhausted receiver, the surface of the mercury in S being acted upon only by the pressure of the rarefied air in the receiver, the weight of the higher column in B A will predominate, and the mercury will fall in it, until the difference of the levels in the two

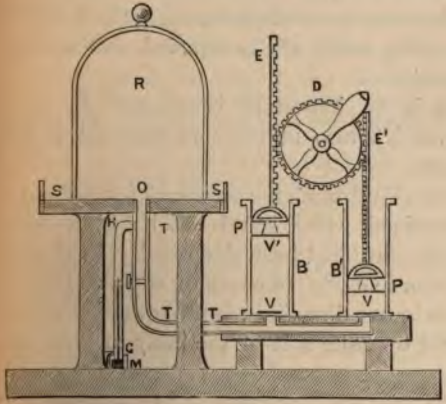

FIG. 33.

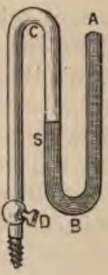

FIG. 34.

legs shall be equal to the pressure of the rarefied air in the receiver.'[1]

201. All the air in the receiver cannot be removed, for when the air has become so rarefied that its elastic force is not sufficient to raise the valve V, the degree of exhaustion attainable has been reached. It is easy to explain the rate at which the air is drawn from the receiver. If we suppose that the volume of the receiver and the pipe together is nine times that of the cylinder, then when the air is diffused through the receiver, the pipe, and the cylinder, that in the cylinder is $\frac{1}{10}$th of the whole. Thus by the up-and-down stroke of the piston we can remove $\frac{1}{10}$th of the air originally in the receiver and the pipe; and therefore we leave $\frac{9}{10}$ths. At every succeeding stroke, $\frac{1}{10}$th of what remained after the preceding stroke will be expelled, and in the same way $\frac{9}{10}$ths will remain.

202. The air-pump is an important instrument in the chemist's laboratory; and it is also in frequent requisition in other branches of science.

203. Some experiments illustrating the pressure of the atmosphere will now be described, which could not be explained satisfactorily until the air-pump had been described. In these experiments it will be seen that the pressure of the air is perceptible only when it acts in one direction; in other words, when there is a vacuum on the opposite side of the body to that on which the air presses.[2]

Experiments Illustrating the Pressure of the Air.

70. On one of the open ends of a strong glass cylinder, a piece of bladder is to be tied so tightly as to be quitè air-tight; the edge of the other open end of the cylinder, which must be ground and well greased, is then placed on the plate of the air-pump and pressed, so as to make the contact air-tight; the pump is then worked to exhaust the air in the cylinder; the bladder becomes first depressed, and finally bursts, accompanied by a loud report caused by the sudden entrance of the air. The exhaustion must be rapid, for if the rarefaction of the air in the cylinder is slow, the bladder distends without bursting, the air passing through its pores.

71. Substitute for the bladder a thin piece of india-rubber; it will be forced, in consequence of its elasticity, into the interior of the vessel. If the top of the cylinder were of thin glass it would be broken.

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[1] Lardner's *Handbook of Natural Philosophy*.

[2] A much simpler and less expensive air-pump than the one described is sufficient for educational purposes.

204. The experiment with the cylinder covered at one end with bladder shows why the pressure of the atmosphere on our bodies, which on a man of middle size is equal to about 16 tons, does not in the least inconvenience us; for as the air in the cylinder counterbalanced the pressure of the external air, and thus prevented the bladder from being burst by the external pressure until the internal air was pumped out. in like manner the air and liquids in the bodies of animals produce a pressure outwards *equal* to the pressure of the external air inwards; the external and internal pressure being equal, no part of the body is, consequently, crushed.

205. The following experiment with the apparatus called the *Magdeburg* hemispheres shows very strikingly the force of the pressure of the air :—

72. The apparatus consists of two hollow brass hemispheres with evenly ground edges, which admit of being brought into air-tight contact when smeared with lard; one of the hemispheres is provided with a stopcock, and its end terminates in a screw by which it can be screwed on the plate of the air-pump, and on the other hemisphere there is a handle. When the hemispheres are joined together, and as long as they remain filled with air, they can be easily separated, as the external pressure of the atmosphere is counterbalanced by the elastic force of the air in the apparatus. When screwed on to the plate of the air-pump for the purpose of pumping out the air, the stopcock is opened; and when a sufficient exhaustion has been effected, it is closed, and the apparatus is unscrewed from the pump-plate; it will now require a powerful force to separate them; and as this is the case in whatever position they are held, it follows that the pressure of the air is transmitted in all directions. The section of the hemispheres employed by the inventor of the apparatus, Otto von Guericke, measured 113 square inches, and they were held together by a force equal to about three-fourths of a ton.

206. We have pointed out in par. 172 a means of proving that it is pressure of the atmosphere that supports the column of mercury in the barometer tube; another and most complete demonstration is shown by placing the instrument under the receiver of the air-pump and exhausting; the column sinks as the pressure diminishes; and when the air is readmitted into the receiver, the column recovers its original height.

73. Place a flask, about half-full of water under the receiver, the mouth of the flask being closed by a cork through which passes a glass tube, which dips into the liquid; when the air in the receiver becomes sufficiently rarefied, the elastic force of the air in the flask causes a jet of water to issue from the tube.

74. Introduce a piece of coke into a vessel of water, and then place the vessel under the receiver; on first exhausting the air an evolution of

mercury in the barometer; in other words, it is a barometer whose Torricellian vacuum is the receiver R. This apparatus has been used with great success in experiments in which a very complete exhaustion was required.

211. We have now to consider more fully than has yet been done in the previous part of the work, the physical properties of gaseous bodies. The first property to be noticed is their *expansibility*, *tension*, or *elastic force*, as this property is indifferently termed.

212. Owing to the almost, if not complete, absence of any attraction between their molecules, they necessarily would be continually tending to occupy a larger space, unless their extension was checked by some boundary, or by some counteracting force, which was equal to their own force of expansion.

213. Liquids exert a pressure on the *sides* of the vessels containing them; but gases exert a pressure in *all* directions upon the surfaces by which they are bounded, and the pressure is equal in all directions. If the resisting medium offer an equal resistance on all sides, the gas, by virtue of this eccentric expansibility will assume a spherical form. Soap-bubbles, and the small bubbles which rise in water, have invariably a globular figure, and show the eccentric expansibility of atmospheric air.

214. If a vessel containing a gas does not enlarge with it, or if the elastic force of the gas exceeds the cohesive force of the solid material in which it is enclosed, then a disruption of the solid takes place, and some of the gas escapes. But if the vessel is subjected to an external pressure, which is equal to the internal pressure, then the vessel suffers no disruption, however weak its sides may be, as we have shown is the case with regard to the human body.

Experiments on the Expansion and the Compression of Gases.

75. Place a bladder fitted with a stopcock, which is closed, the bladder being half-full of air, under the receiver of the air-pump; pump the air out of the receiver, and immediately the bladder will distend, and on readmitting the air into the receiver it will immediately collapse to its original bulk.

76. Press the rod of the piston of the *pneumatic* syringe, an apparatus which consists of a stout glass tube closed at one end, and provided with a tight-fitting solid piston; it will of course move down the tube, the effect being that the air is compressed into a smaller volume; when

re is removed, the piston is pushed upwards by the elastic force air to its original position, the air regaining its original volume.

To one of the necks of a two-necked Woulfe's bottle containing water, as represented in fig. [illegible], a bladder filled with air is tied on; in the other neck a glass tube is fitted air-tight by means of a [illegible] cork, the lower end reaching almost to the bottom of the [illegible]. In compressing the bladder, the water, owing to the compression air, will ascend in the tube.

15. We have noticed the great difference there is between [illegible] in the gaseous state and in its two other states of aggre[illegible] as regards its compressibility, and we have explained [illegible] shown by experiment that air is capable not only of [illegible]pression, but that with its [illegible]pression its density and [illegible] force increase. The [illegible] that expresses the com[illegible]sibility of gaseous matter discovered by the Hon. [illegible] Boyle in England in [illegible] and by Mariotte in [illegible] in 1676. The law [illegible] the names of both philo[illegible]. It may be expressed [illegible]lows:

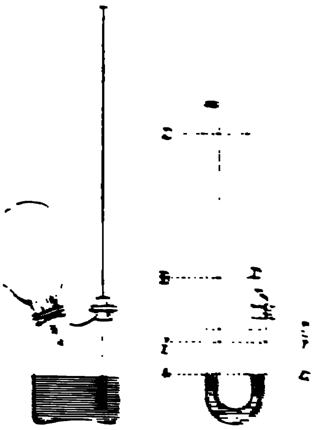

Fig. [illegible]

Fig. [illegible]

1. *The elastic force of a [illegible] density [illegible] the [illegible] the volume.*

2. *The density of a [illegible] the pressure, and [illegible] the volume.*

3. *And the volume of a [illegible] is [illegible] pressure.*

16. It may be proved by the following simple experiment. A syphon-shaped tube [illegible] of uniform bore is employed [illegible] longer leg is open at the top, the end of the shorter one is [illegible] hermetically sealed, or it is furnished with a stopcock as [illegible] in the figure. The stopcock being opened, a quantity [illegible] mercury is poured into the open end. The surfaces of the [illegible] will of course stand at the same level in both. The two columns of air [illegible] and [illegible] sustain a pressure [illegible] the weight of a column of air [illegible] from A and [illegible]

to the top of the atmosphere. If we now close the stopcock, D, the effect of the weight of the whole atmosphere above that point is cut off, so that the surface, *a*, can sustain no pressure arising from the weight of the atmosphere. Still the level of the mercury remains the same, because the elasticity of the column of air, *a* D, is precisely equal to the weight of the whole column before this small length was cut off. The surface, A, is still pressed by the whole atmospheric column, and thus we see that these two different properties of the atmosphere,[1] its *elasticity* and its *weight*, exactly counterbalance each other.

217. 'We have learned that the atmospheric pressure, under ordinary circumstances, is equal to 15 lbs. on the square inch, or to a column of mercury 30 inches high. It is evident therefore, that the atmospheric pressure acting on A is the same as would be produced by a column of mercury, 30 inches high, resting on the surface of A. So, also, the force with which the air confined in *a* D presses by its elasticity on the surface, *a*, is also equal to a column of mercury 30 inches high. The pressure of the atmosphere acting on the surface, A, is transmitted by the mercury to the surface *a*, and balances the elastic force of the isolated column, *a* D.

218. 'If we now pour an additional quantity of mercury into the open end of the tube at C, an increased pressure, arising from the weight of the metal, will be transmitted to the surface, *a*, and will prevail over the elasticity of the confined air; the surface, *a*, will therefore rise towards D, compressing the air into a smaller space. On continuing to pour in mercury, until the surface, *a*, rises to *b*, or half-way between *a* and D—that is, until the confined air is compressed into exactly half its former limits—it will be found, on drawing a horizontal line from the surface *b* to the opposite *b'* in the longer limb, that the column of mercury, *b'* B, measures 30 inches, the weight of which is equal to the atmospheric pressure. The force with which the surface, *b*, is pressed upwards towards D, is therefore equal to *two atmospheres*, or double the force with which *a* was pressed upwards towards D. Hence it appears that the elasticity of the confined column of air, *b* D, is double its former elasticity when filling the space *a* D; so that when the air is compressed into half its volume, its elasticity is doubled. If we again pour mercury into the tube at C, until the air enclosed in the shorter limb be reduced to a third of its bulk, as at *c* D, the compressing force will be equal to three times the atmospheric pressure. The height of the compressing column of mercury would reach to C, namely 60 inches above the level, *e*. If we still add mercury, until the column rise to the height of 90 inches above its level in the short limb, the elastic force of the confined air would be four times greater than at first, and it would be compressed to the bulk of one-fourth its original volume. The experiment proves that *the tension or elastic force of gases varies in exactly the same proportion as their density.*

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[1] This applies not only to air but to all gaseous bodies.

219. 'As in a state of rest the tension everywhere exactly balances the pressure, the two terms are frequently interchangeable, although it is usual to estimate pressure as so many pounds on a square inch, and to measure the tension by the height of the column of mercury which it is capable of sustaining. Either of these measures, however, can always be easily reduced to the other.'[1]

220. When a gas in a confined space is heated, its tension increases by $\frac{1}{273}$ its value at 0° C. for every degree Centigrade through which its temperature is raised. This is part of Charles's law and may be thus expressed:—*The tension of a constant volume of gas varies directly as the absolute temperature.*

221.—The rate of expansion of gases at increased temperatures has been stated in general terms. We must now enter more fully into the subject. It was there stated that the rate of expansion of all gases is equal and uniform at all degrees of heat. Although this is sufficiently accurate for all chemical requirements, the coefficient of expansion is not rigidly uniform for all gases.[2] A volume of any gas at the melting-point of ice increases in bulk, the pressure remaining constant during the operation, to 1·3665 volumes when raised to the temperature of boiling water. As on the Centigrade scale the interval between the melting-point of ice and the boiling-point of water is, as we have learned, divided into 100 degrees, it follows that any gas which measured one volume at the melting-point of ice would increase 0·003665 in bulk, for each degree it increased in temperature.

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[1] Tomlinson's *Pneumatics.*

[2] The more readily condensable gases display irregularities, especially near their points of solidification and liquefaction. Regnault and Magnus have published independent and elaborate investigations on the expansion various gases undergo by the application of heat. According to their experiments the coefficient of expansion is not rigidly uniform for all gases, the expansion being

222. As the interval, between the melting-point of ice and the boiling-point of water on the Fahrenheit scale, is divided into 180 degrees; it results that, for each degree F. that the gas increases in temperature, it must increase in bulk $\frac{1}{491}$. Knowing the coefficient of expansion, we can compute the volume a gas would occupy at some other temperature than that at which it was measured; we will explain how these calculations are made, and give a few exercises for the student to work out.

223. *To calculate the volume of a gas from a lower to a higher temperature on the Centigrade scale*, we multiply the number by 0·003665, which is the amount we have learned one volume of any gas expands for each degree C., and to the product thus obtained we add 1 for the volume unit; we then multiply the number thus obtained, by the number of volume-units the gas measured at the lower temperature; this will give us the volume the gas we have measured would occupy at the higher temperature. Suppose, for the sake of illustration, we wanted

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the greatest for those which are the most readily condensable; whilst the least readily condensable exhibit scarcely any appreciable difference.

The following table contains a summary of Regnault and Magnus's experimental results.

Expansion of Gases by Heat.

1000 parts, at 32° F., become	At 212° F.	
	Regnault	Magnus
Hydrogen	1366·13	1365·659
Carbon monoxide	1366·88	
Atmospheric air	1367·06	1366·508
Nitrogen	1366·82	
Hydrochloric acid	1368·12	
Carbon dioxide	1370·99	1369·087
Protoxide of nitrogen	1371·95	
Cyanogen	1387·67	
Sulphurous acid	1390·28	1385·618

to know what volume a gas measured at 0° C. would occupy at 30° C., we multiply 0·003665 by 30 and add one for the volume-unit, which gives 1·01099; we then multiply the number we have obtained by the number of volume-units the gas measured at the lower temperature; this will give us the volume the gas we have measured would occupy at the higher temperature. In our example the lower temperature is 0° C., and the higher temperature is 30° C., and suppose the gas measured 20 cubic inches at 0° C. we should have to multiply 1·01099 by 20; the product would be the volume the 20 cubic inches the gas measured at 0° C., would occupy at 30° C.; thus,—$1·01099 \times 20 = 20·2198$, the number of cubic inches the gas would measure at 30° C.

The following would be the general formula for calculating the volume of a gas from a lower to a higher temperature, according to the Centigrade scale, v denoting the known volume, v_1 the unknown volume, and t the temperature :—

$$v_1 = v \times (1 + {·}003665 \times t).$$

224. *To reduce the volume of a gas from a higher to a lower temperature on the Centigrade scale*, the number of volume-units the gas measures at the higher temperature must be divided by the number obtained by multiplying 0·003665 by the thermometric degree at which the gas is measured, and adding to the product 1 for the volume-unit; the quotient will be the volume the gas will occupy at the lower temperature. As an example, we will reverse our former illustration. What volume will 20·2198 cubic inches of gas at 30° C., occupy at 0° C.?

$$\frac{20·2198}{1 + 0·003665 \times 30} = 20 \text{ cubic inches.}$$

The following is a general formula for reducing the volume of a gas from a higher to a lower temperature, according to the Centigrade scale :—

$$v_1 = \frac{v}{1 + {·}003665 \times t}.$$

225. *The volume of a gas from a lower to a higher temperature on the Fahrenheit scale may be calculated thus*: 491 volumes of a gas increase one volume for every increase in degree on the F. scale; therefore 491 volumes at 32° F. would measure 509 volumes at 50° F., because there is a difference of 18° between 32° and 50°; therefore if we wished, for example, to know what volume 20 cubic inches of gas at 32° F. would measure at 50° F., we make the following proportion sum:—

$$491 : 509 :: 20 : 20{\cdot}733.$$

The general formula would be (t_1 denoting the temperature of the unknown volume):—

$$v_1 = \frac{v \times 491 + (t_1 - t)}{491}$$

If, instead of calculating the volume of a gas from a lower to a higher temperature, we were by this plan to reduce the volume from a higher to a lower temperature, the calculation would be as follows, reversing our former illustration, viz., What volume will 20·733 cubic inches of gas, at 50° F., occupy at 32° F.?

$$509 : 491 :: 20{\cdot}733 : x = 20 \text{ cubic inches.}$$

The general formula would be:—

$$v_1 = \frac{v \times 491}{491 + (t - t_1)}$$

Exercises on the computation of the volume of gases at different temperatures.

78. A gas measures 20 cubic inches at 50° F. What would it measure at 60° F.?

79. A gas measures 15 cubic inches at 90° F. What would it measure at 60° F.?

80. A gas measures 7·3 cubic inches at 180° F. What would it measure at 60° F.?

81. A gas measures 100·3 cubic inches at 40° F. What would it measure at 60° F.?

82. A gas measures 10·7 cubic inches at 320° F. What would it measure at 80° F.?

83. A gas measures 15 cubic inches at 0° C. What would it measure at 60° C.?

84. A gas measures 40·1 cubic inches at 100° C. What would it measure at 0° C.?

85. A gas measures 3·5 cubic inches at 0° C. What would it measure at 80° C.?

86. What would a decilitre of gas at 0° C. measure at 80° C.?

87. What would 7 cubic centimetres at 24° C. measure at 0° C.?

88. A room of 2000 cubic feet capacity, is heated from 60° to 80° F.

How much air is expelled from the room, using Regnault's coefficient of expansion, viz. $\frac{1}{491}$ of the bulk at 32° for each degree of Fahrenheit's thermometer?

226. As hot air is lighter than cold air, air as it becomes heated rises and accumulates in the upper part of a room. This is easily demonstrated by partially opening the door of a warm room, and holding a lighted candle near the top; if the window be not open, the flame will be forced in the outward direction by the current of heated air which is escaping from the room. Midway down, the flame will not be affected either inwards or outwards, it will be stationary; while near the floor it will be blown into the room by reason of the cold air entering at the bottom.

227. By the respiration of human beings and other animals, and by the burning of substances for the production of heat and light, the oxygen of the air is being constantly removed, and is replaced by carbon dioxide, which is emitted from both these sources; and this gas is inimical to animal life, and is incapable of supporting combustion. The air of a room where people are assembled requires to be constantly renewed, whether combustible substances are being burnt or not. The quantity of carbon dioxide evolved by the lungs and skin of one person 'amounts to about 0·7 cubic foot per hour. In order that it may be breathed again without inconvenience, this should be distributed through at least 140 cubic feet of fresh air, or a space measuring 5·2 feet each way'; hence the necessity for a constant supply of fresh air by *ventilation* even when no gas, candles, &c., are burnt, but when combustibles are burnt, the necessity is increased enormously. The contaminated air proceeding from both sources—respiration and combustion—being of higher temperature than the surrounding air ascends, colder air which is fresh, taking its place; it requires, therefore, only proper ventilation of rooms for this contaminated air to be carried away.

228. The ventilation of rooms, the student will see from what has been stated, is an important art, and is dependent on the currents produced in air by differences of temperature. The draught of a fire is simply caused by the ascent of the heated air in the chimney, and the ascent is quickened by reason of the chimney keeping the heated air together ; and the longer the chimney the more powerful the draught ; this is the reason why factory chimneys are made of such a length ; it is not, as many suppose, for solely carrying the smoke high up in the air. As tall chimneys quicken the draught, chimneys of short length have only a moderate draught ; hence the chimneys of cottages, and the upper rooms of houses are more liable to smoke than those that are longer, as the draught is proportioned to the length of the chimney. If the section of the chimney is larger than is necessary to allow an exit for the products of combustion, descending as well as ascending currents are produced in it, which causes it to smoke. In *open fireplaces*, a large quantity of the colder air of the room enters the chimney above the fire, and mixes with air heated by the fire ; this mixture ascends more slowly than if hot air alone had entered the chimney. By excluding a portion of the cold air by placing a plate of metal or even a piece of paper across the opening, the fire begins to burn more quickly, by reason of the draught being increased, owing to the exclusion of the cold air. In *close fireplaces*, as those of steam-engines, when the furnace door is shut, no air can enter the chimney but directly through the fire ; a powerful draught is consequently produced. 'Rumford was the first to attempt rational improvements in chimneys. He reduced the opening of the chimney and the depth of the fireplace, and added polished plates inclined at an angle which serve both to guide the air to the fire, and to reflect heat into the room.' An open fireplace, however improved, will always be an imperfect and costly mode of heating ; it is estimated that only 13 per cent. of the total heat is made available ; but although an imperfect and costly mode of heating, it ensures *good ventilation.* The fresh air to supply that carried up the chimney is usually supplied through the crevices of the doors and windows ; if no fresh air were admitted into the room, there would be produced a downward current of air in the chimney, and as a consequence the chimney would smoke.

Exercise on Ventilation.

89. A chimney, 4 feet in length, is filled with air of the average temperature of 104° F., the outer air being 44·5° F. What height would be the column of the outer air, which the air in the chimney would balance?

And what would be the ascensional force of air in the chimney owing to the difference in weight of equal columns of the inner and outer air?

229. In par. 4 it is stated that matter in one state of aggregation can be made to pass into one or both of the other two states of aggregation, by the application of heat, cold, or pressure, or by a combination of pressure and one of the other two agents. All known gaseous bodies have now been liquefied; some are easily liquefied, either by pressure or cold; others require intense pressure combined with extreme cold; and some have been solidified.

230. Faraday, who was one of the earliest to liquefy gases, adopted a very simple method for liquefying the gases that are most readily liquefied; a substance, containing the gas to be liquefied as a constituent, he placed in one of the limbs of a strong narrow bent tube (fig. 38), the open end of which was then sealed up. The gas was disengaged from the substance containing it by heating that end of the tube, and immersing the other limb in cold water; the gas accumulating in the tube became liquefied by its own pressure. With less easily liquefied gases he had to employ cold with the pressure. But with this apparatus he was unable to liquefy many gases: these have since yielded to more intense pressures and more extreme cold than he was able to employ; for the liquefaction of hydrogen and oxygen, for example, it requires for the liquefaction of the former a pressure of 450 atmospheres, coupled with a temperature of $-140°$ C.; and for the latter 22·5 atmospheres, and a temperature of $-136°$ C.

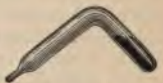

FIG. 38.

231. When the pressure is removed the liquefied gas volatilises, and its volatilisation is, of course, attended with a great diminution of heat, so much so that in some cases a portion of the liquefied gas is solidified; in this way the gas, carbon dioxide, was long ago obtained in the solid state; and very recently oxygen had also been solidified by like means.

232. From what has been stated about the liquefaction of gases, the student will no doubt generalise that some bodies that are gaseous in our climate would assume the liquid state in the Arctic regions; just as some bodies that are in the liquid state in our climate would there be in the solid state, and he will have generalised correctly; he may also, probably, further generalise, and the generalisation will be correct, that some bodies that exist in the liquid state in our climate become gaseous in tropical countries. The vapourisation of liquids will form the subject-matter of the next chapter.

Examination Questions.

90. Explain the distinction between mass and weight.

91. Explain the principle of the syphon, and state the limitation of its action.

92. Describe the barometer, and explain how it may be used for determining the difference of the altitudes of two stations.

93. A body when placed successively in the two pans of a balance, weighs in one pan 63 grammes, and in the other 65; calculate its true weight.

94. Give your reasons for preferring an air thermometer as the best means for measuring temperature.

95. Suppose you wish to verify the accuracy of a Centigrade thermometer at its freezing and boiling points, describe in detail how you would proceed to do so; which point would you commence with first? Give a reason for your reply.

96. If ice at the freezing-point is dissolved in ten times its weight of water at 20° C., what will be the temperature of the mixture?

97. A small cube of wood is placed floating in a glass of water, the glass is then placed under the receiver of an air-pump, which is then partially exhausted: will the wood rise or sink? Explain clearly your answer.

98. State the law of dilatation of gases for changes of temperature.

99. 40 cubic centimètres of atmospheric air are enclosed in a tube over mercury, the height of the mercury in the tube above the level in the trough outside being 50 cubic centimètres. The tube is depressed, until the mercury in the tube above the level in the trough is equal to 30 centimètres. What is now the volume of air? The temperature (15° C.), and the height of the barometer (76 centimètres), have not changed during the observations.

100. It is found that a kilogramme of water at 100° C., mixed with a kilogramme of snow without loss of heat, gives 2 kilogrammes of water at the temperature of 10° C. Show how to find from this the latent heat of water.

101. As gravity acts on all substances with the same intensity, why do bodies fall through the air with unequal velocities?

102. A certain quantity of air at atmospheric pressure has a volume of 2 cubic feet, the temperature being 55° F. What does the volume of air become when the pressure is increased by $\frac{1}{20}$, the temperature meanwhile remaining the same?

103. Can you assign any reason why two chimneys should not communicate?

104. If only 13 per cent. of the total heat yielded by coal is obtained by burning it in an open fireplace, what becomes of the rest of the heat?

105. Explain scientifically why a chimney ought to be at least of a certain height.

106. State precisely, and illustrate by numerical example, the laws which give us the relations between the temperature, the pressure, and the volume of a given gas.

107. 500 cubic centimètres of oxygen are measured when the tem- ature is 20° C. ; the temperature is subsequently raised to 40° C., the ssure meanwhile remaining invariable. What is the volume of oxygen the latter temperature? (The coefficient of expansion of oxygen is).

108. Air is allowed to expand to double its original volume ; how ch must its temperature be raised that its pressure may remain un- nged ?

109. Let the volume of a gas be 5 litres at the pressure of 290 milli- tres of mercury, and at 10° C. What will be its volume at the pressure 50 millimètres and at 72° C. ? Use the coefficient 0·003665.

CHAPTER IV.

APORATION—EBULLITION—DISTILLATION—TENSION OF VAPOURS—CORRECTION OF GASES FOR PRESSURE—EXERCISES—CORRECTION OF GASES FOR TEMPERATURE, PRESSURE, AND TENSION OF AQUEOUS VAPOUR—EXERCISES—QUESTIONS.

233. SOLID and fluid bodies, in the gaseous state, are called pours, and their conversion into this state is called evapora- n, or vaporisation. Vapours, of which steam is the most niliar to us, are light, expansible, and generally colourless ses, resembling air, and other gaseous bodies ; but vapours more readily condensed by a change of a temperature than dies which are gaseous under the ordinary conditions both temperature and pressure.

234. Vapours, when they are separated from the liquids ich have yielded them, obey the same laws as gases ; but if e liquid is present, heat not only expands the vapour, but so *increases its volume* by the addition of a new quantity of pour.

Example : Steam, when heated by itself, apart from the uid which produced it, does not possess a greater elasticity an an equal bulk of air confined and heated to the same de- ee, and may be heated to the temperature at which the con- ining vessel becomes red-hot, without acquiring great elastic rce ; but if water be present, then more and more steam ntinues to rise, adding its elastic force to that of the vapour eviously existing, so that the pressure becomes excessive.

235. The passage from the liquid state to the state of vapour *is distinguished from* the passage of the solid state to the liquid state by the important fact that, whilst the liquefaction of each solid is produced *at one* temperature, the vaporisation of each fluid occurs at a *variety* of temperatures.

236. When a liquid vaporises, with or without the application of heat, and the elastic force of the vapour is not equal to the pressure of the atmosphere, the process is called evaporation, and the liquid is said to evaporate.

237. When a liquid vaporises, with or without the application of heat, and the elastic force of its vapour is equal to the pressure of the atmosphere, the vapour produces a bubbling or boiling, and the liquid is said to boil; and the temperature at which this takes place is called the boiling-point of the liquid. In evaporation, the vapour is formed only from the *surface* liquid; in boiling, the vapour is formed from the *whole mass* of the liquid.

238. It will be evident, from what has been stated, that if the pressure of the atmosphere be diminished, either artificially or by ascending to greater heights in the air, as the tops of lofty mountains, liquids will, on account of the diminished pressure boil at lower temperatures;[1] and if the pressure be increased, either artificially, or by descending into mines, liquids will, on account of the increased pressure, boil at higher temperatures: in other words, *the boiling-point of liquids varies with the pressure.* Nevertheless, the temperature at which each liquid boils is *quite constant under the same circumstances*, and the boiling-point often becomes a physical character of great importance in distinguishing liquids which much resemble each other; as different liquids, under the same circumstances, boil at very different temperatures. Ether, for example, boils at 94°·8° F., whilst mercury boils, under the same circumstances, at 662° F.

239. The following very simple and beautiful experiment proves that, with the diminution of pressure, liquids boil at lower temperatures. A little water is made to boil in a flask until all the air is expelled, and the steam issues from the neck. A tightly fitting cork is then inserted, and the lamp at the same moment withdrawn. By plunging the flask into cold water the *boiling is renewed*, because a partial vacuum is produced by the cold water condensing the steam in the flask. The *boiling ceases* on plunging the flask into hot water, because the steam ceases to be condensed, and its pressure stops the boiling.

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[1] We have already noticed that the heights of mountains can be measured by means of the barometer; they can also be measured by ascertaining, by means of a very delicate thermometer, at what temperature water boils. A difference of about 1° F. is occasioned by an ascent of about 535 feet.

240. In a Papin's digester, which is a tight and strong kettle with a safety valve, water may be raised to 300° F., or 400° F., without boiling, but the instant that the great pressure is removed, the boiling commences with great violence.

241. There are other circumstances besides pressure which influence the temperature at which liquids boil. The presence of air in a solution has been found to assist the evolution of vapour. Water deprived of air may be raised many degrees above its boiling-point, and it will then be suddenly converted into steam with explosive violence. 'The nature of the vessel, or rather the state of its surface, also exercises an influence on the boiling-point, and this to a much greater extent than was formerly supposed. It has long been noticed that in a metallic vessel water boils, under the same circumstances of pressure, at a temperature one or two degrees below that at which ebullition takes place in glass; but it has been shown that, by particular management, a much greater difference can be observed. If two similar glass flasks be taken, the one coated on the inside with a film of shellac, and the other completely cleansed by hot sulphuric acid, water heated over a lamp in the first will boil at 211° F., whilst in the second it will often rise to 221° F., or even higher. A momentary burst of vapour then ensues and the thermometer sinks a few degrees, after which it rises again. In this state, the introduction of a few metallic filings, or angular fragments of any kind, occasions a lively disengagement of vapour, while the temperature sinks to 212° F., and then remains stationary. These remarkable effects must be attributed to an attraction between the surface of the vessel and the liquid. The solution of solids in a liquid, on account of the adhesion between them, also causes the liquid to boil at a higher temperature.

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242. An opinion formerly prevailed that not only bodies which are vapours at a moderate temperature, but even substances, as the earths and some of the metals, which cannot be volatilised in the fire, emit vapours at all temperatures, even the lowest. This has been shown by Faraday to be an erroneous opinion. He has proved that there is a temperature below which volatilisation ceases—a temperature which varies for different substances. For mercury, the limit is about 40° F., for sulphuric acid the limit is much higher, since the acid undergoes no sensible evaporation at ordinary atmospheric temperatures. The cohesive force between the particles of a liquid is greater probably than the repellant power between the particles at the temperature at which the liquid ceases to evaporate, and it is this which puts a stop to the evaporation.

243. Some solid bodies, such as camphor and arsenic, pass at once into vapour without being liquefied ; but most solids become liquids before they are converted into vapour. Even ice and snow can evaporate without first liquefying. It is owing to this evaporation that patches of snow and tufts of ice are observed gradually to disappear even during the continuance of a severe frost.

244. On the application of heat to a liquid, the temperature of the liquid continues to increase until the boiling-point is reached ; but the moment ebullition commences, the temperature of the liquid remains stationary, however intense the heat applied may be, and however long continued. This fact is of importance in domestic economy, particularly in cookery, and attention to it would save much fuel. Soups, &c., made to boil in a gentle way, by the application of a moderate heat, are just as hot as when they are made to boil on a strong fire with the greatest violence.

245. Although liquids are converted into vapours at and below the ordinary temperature of the air, they nevertheless require for this conversion a large amount of heat, which becomes latent in the vapour. If, for instance, a few drops of ether be allowed to fall on the hand, the ether evaporates, and a sensation of cold is experienced, because the ether requires for its vaporisation a large amount of heat, which it obtains from the hand, hence the cold. Water may even be frozen by the rapid absorption of heat occasioned by its own evaporation. The porous water-jars which are used in warm countries to keep liquids cool owe their efficacy to the latent heat of vapours.

246. The sensible heat of vapours is the same as that of the liquids from which they are produced, therefore a vapour produced from a liquid which has the temperature of 60° F. will also have that temperature; but if the liquid had a temperature of 212° F., the vapour would also have a temperature of 212° F.

247. Equal weights of different liquids require very different amounts of heat to convert them into vapour. But a much larger quantity of heat is necessary to convert fluids into vapours than solids into fluids. As an example, observe the quantity of heat necessary to convert ice into water, and water into steam: 1 lb. of ice at 32° F. requires for its conversion into 1 lb. of water at 32° F., 142·65° F. of heat; in other words there is as much heat absorbed in the liquefaction of 1 lb. of ice as would raise 142·65 lbs. of water through 1° F.; but 1 lb. of water at 212° F. requires for its conversion into 1 lb. of steam at 212° F., 966·6° F.; so that 966·6° F. of heat have become latent in converting water into the vaporous state; the latent heat of steam is therefore 966·6° F.; consequently, the conversion of 1 lb. of steam at 212° F. into water at 212° F. would evolve as much heat as would raise 966·6 lbs. of water 1° F. If the heat necessary for the vaporous state is abstracted, the vapour is converted into the liquid state. The well-known process of distillation is an application of this principle.

248. Equal bulks of different liquids produce very different volumes of vapour. Water furnishes, bulk for bulk, a much larger amount than any other liquid, a cubic inch of water at 212° F. expanding to nearly a cubic foot of steam at 212°, or to state it accurately, to 1,696 times its volume. The following table shows the volume of vapour which is furnished by a cubic inch of four different liquids, at their respective boiling-points. Equal volumes of different vapours, taken at the boiling-points of their respective liquids, consequently possess very different weights, as is shown by the last column of the table:—

1 cubic inch of each liquid at 60° F., yields in the case of	Cubic inches of vapour at its boiling-point	Boiling-point °F.	Weight in grains of 100 cubic inches of each vapour at its boiling-point
Water	1,696	212	14·93
Alcohol	528	173	40·49
Ether	298	95	64·71
Oil of Turpentine .	193	314	117·71

249. Before proceeding farther with respect to the elastic force of vapours, it is necessary to describe the mode of determining the boiling-point of liquids, and their distillation, as these are operations in frequent use in chemical laboratories. The size of the vessel in both operations is determined by the quantity of liquid at the command of the operator.

250. We have already noticed that the temperature at which each liquid boils is quite constant under the same circumstances. We have also learned that other circumstances besides pressure influence the temperature at which liquids boil; but these circumstances do not affect the temperature of the vapour of the boiling liquid ; this at once adjusts itself to the pressure of the atmosphere, and is always constant for the same liquid under the same pressure. In determining, therefore, the boiling-point of a liquid we do not allow the bulb of the thermometer to dip into the liquid, but we bring it almost into immediate contact, so that it may be completely surrounded by the vapour.

251. As the boiling-point of any body of definite composition, and perfectly pure, is stationary during the entire time it is kept boiling, whilst the boiling-point of a mixture of substances keeps constantly increasing as the more volatile substances evaporate, the boiling-point becomes the most important physical character for determining the identity and purity of a volatile liquid.

252. When we have only a small quantity of liquid to operate upon we may employ a test-tube fitted with a cork; the cork must be pierced with two holes, in one of which is inserted the thermometer, and in the other the exit-tube. If we have a large quantity of liquid at our command, we may employ a tubulated retort ; the tubulure is fitted with a cork pierced with one hole, in which the thermometer is inserted.

If we do not wish to lose any of the liquid, we condense the vapour by connecting the exit-tube from the test-tube, or the beak of the retort, with a Liebig's condenser (see fig. 40). The liquid is allowed to boil until the greater part of it has distilled over: we note the pressure of the barometer, and we also note the temperature during the whole time.

253. The temperature indicated by the thermometer is less than that of the true boiling-point, if only part of the stem of the thermometer is exposed to the vapour of the liquid. For a thermometer to indicate the true temperature of vapour or any other medium, it is necessary to immerse both the bulb and stem in the medium whose temperature is to be measured; for if the stem of the thermometer is exposed to a lower temperature than the bulb, the whole of the mercury will not be fully expanded, and the thermometer will indicate too low a temperature: this difference between the indicated and the true temperature becomes important when the temperature of the medium greatly exceeds that of the air.

254. We can adopt one of two methods for obtaining the exact temperature at which a liquid boils: we may either have, by a simple contrivance, the stem of the thermometer immersed in the vapour of the boiling liquid; or we may by calculation make the necessary correction. We will describe both methods:—

1. The following simple apparatus for the exact determination of boiling-points has been devised by Berthelot. It consists of a long-necked glass flask (fig. 39), having a capacity of 100–200 c.c. The neck is enclosed in a wide glass tube, closed with two perforated corks, and passes through the lower cork; and the tube of a thermometer passes through both; its bulb and stem are

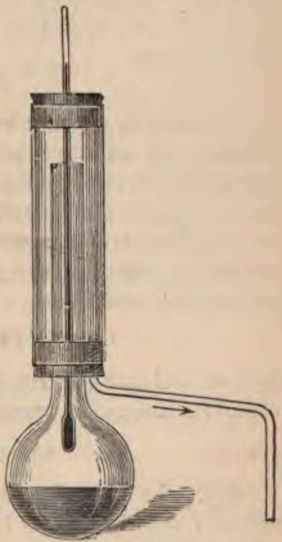

FIG. 39.

thus immersed in the vapour of the boiling liquid. The lower cork is also traversed by a bent tube, which serves to carry off the vapour to a condenser, and thus prevent any excess of pressure in the apparatus. Due attention must be paid, in taking the observations, to the barometric pressure at the time.

2. For making the correction by calculation it is necessary to know the mean temperature of the exterior portion of the column of mercury exposed to the air, compared with that part below the cork A, surrounded by the vapour; for this purpose we fix to the stem of the first a second thermometer, in such a way that its reservoir is about the middle of the column of mercury which is above the distillatory apparatus.

If we represent by T° the temperature indicated by the thermometer in the vapour, by t° the mean temperature, and by N° the number of degrees which the exterior mercury column measures; lastly by δ, the coefficient of the apparent expansion of mercury in glass ($\delta = 0{\cdot}000154$ for 1° C), the quantity to be added to T° for obtaining the true boiling-point becomes

$$N.\ (T-t)\ \delta.$$

As an example, suppose we have found $171{\cdot}3^\circ$ C.$=T$ for the boiling-point of the liquid, with a thermometer having 25 degrees above zero in the distillatory apparatus; in this case, $171{\cdot}3^\circ - 25$, or $146{\cdot}3^\circ = N$. If the exterior thermometer indicates $43^\circ = t$ for the mean temperature of this portion of the column of mercury, $171{\cdot}3 - 43^\circ$, or $128{\cdot}3 = (T-t)$. The correction becomes then—

$$146{\cdot}3^\circ \times 128{\cdot}3 \times 0{\cdot}000154 = 2{\cdot}9^\circ,$$

and we find that the true boiling-point of the liquid under the observed barometric pressure is

$$171{\cdot}3^\circ + 2{\cdot}9^\circ = 174{\cdot}2.$$

255. The operation by which a liquid is converted into vapour; and the vapour is re-converted into the liquid state, is termed *distillation*. In this form of distillation, which we shall term *ordinary* distillation, there is no alteration in the constitution of the substances which distil over; but there is a form of distillation, which is termed *destructive* distillation, in which the original substance is *destroyed*, and gaseous and liquid bodies are generated from the destruction, and they distil over; the manufacture of coal-gas is a familiar example of destructive distillation. We shall at present explain only ordinary distillation.

256. We can by means of ordinary distillation separate liquids of different volatilities; and we can separate liquids from the solid bodies which may be dissolved in them. The chemist in most of his operations, has to employ pure water, he has therefore to distil the well or river water to free the water from the solid substances these waters always contain; the pure water thus obtained generally goes by the name '*distilled* water.' When fresh water is not to be had on board ship, or in the neighbourhood of the sea, the sea-water is distilled to obtain this necessary liquid.

257. The *still*, as the apparatus is termed which is used on the large scale, consists of three parts, viz. the body which holds the liquid to be distilled (the body is placed over a furnace): from the body proceeds a tube for the vapour to escape by, this tube terminates in a spiral pipe of tube called the *worm*; the worm is in a kind of cistern which is being constantly supplied with cold water during the distillation. The vapour is condensed in the worm, owing to the worm being always kept cold during the operation by the cold water which surrounds it externally. The cold water is conducted to the bottom of the *refrigerator*, as the cistern containing the worm is called, thus displacing the heated water, which, on account of its less specific gravity, rises to the upper part of the refrigerator, and flows off.

258. The arrangement of the distillatory apparatus employed in chemical laboratories varies with the kind of liquid to be distilled; a moderate or a very high temperature, for example, may be required for the conversion of the liquid into vapour as the range of temperature at which different liquids boil and are converted into vapour is very extensive; in like manner the temperature at which the vapours of different liquids are condensed varies exceedingly: some are condensed at such a moderate temperature that no special condenser is required; on the other hand some require an exceedingly low temperature to effect their condensation. In such a work as this, it would be out of place to describe all the various modifications of the apparatus required for the distillation of different liquids; we shall therefore describe the most usual form of the distillatory apparatus.

259. A retort or flask is employed to hold the liquid to be distilled; a stoppered retort (A) is the holder shown in fig. 40. The condenser employed is called, after the name of the inventor, Liebig's condenser (*b*); it consists of an outer metallic or glass tube; into this is fitted, by means of perforated corks, a glass tube, *c*, about 30 inches in length; into this tube the neck of the retort is fitted. Into the space between outer and inner tubes of the condenser a constant stream of cold water flows from the vessel *d*; the heated water running out through the tube *e*. The condensed vapour passes into the receiver *f*. When the liquid has been placed in the retort, and the apparatus is all arranged, and the space between the two tubes in the condenser has been filled with cold water, heat is gradually applied to the retort by means of a gas-burner, as shown in the figure.

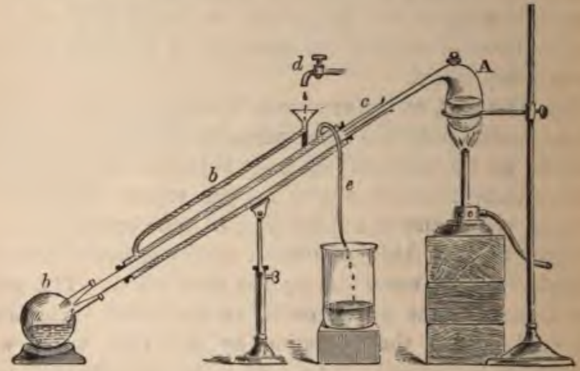

FIG. 40.

260. By a single distillation, carried on in this manner, a liquid can be perfectly separated from one or many non-volatile substances; but two or more volatile liquids cannot be separated by a single distillation, owing to liquids evaporating below their boiling-points; the two liquids pass over together but in different proportions, the more volatile passing over in the largest proportion in the first portion, and the less volatile in the last portions, of the *distillate*; and of course the greater the difference in the boiling-points of the liquids the more perfect the separation effected. But in order to obtain a perfect separation more distillations than *one* have to be made, and portions of the distillate have to be collected separately, the receiver being changed at every increase of 5° C., as the boiling-point of a mixture of liquids is not *stationary*, as the temperature keeps gradually increasing. And to facilitate the separation, a modification of the distilling vessel has to be made; this kind of distillation is called *fractional* distillation, which we will briefly describe.

261. The mode for facilitating the separation of different liquids consists in causing the less volatile of the liquids to condense in the upper part of the distilling apparatus, and flow back into the flask containing the liquid. The best mode yet devised for effecting this condensation is to insert into the mouth of the flask, by means of a perforated cork, a vertical tube represented in fig. 41, having several bulbs blown upon it, and a side tube to assist in conveying back the condensed liquid into the flask. In the mouth of the vertical tube a thermometer is inserted; its bulb is passed down into the uppermost bulb on the tube. To effect the separation of different liquids, even with this modification, the receiver must be changed at every 5° C. And each fraction thus obtained must be submitted to another fractionation until each fraction has a constant, or almost constant, boiling-point.

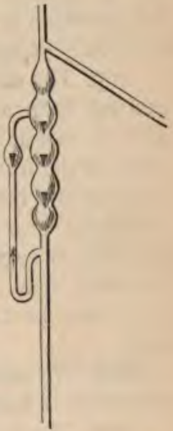
FIG. 41.

262. Wanklyn has shown that we cannot, even by fractional distillation, laborious as the operation is in many cases, separate different liquids, if the *vapour-tensions* and *vapour-densities* are *proportional* to one another, for in such cases the mixture distils over unchanged.

211. Methods for dispensing with frequent and costly [illegible] were employed in the arts long before any simil[illegible] operations. One of [illegible] methods is that adopted in [illegible] beyond the scope of this work [illegible]

[illegible] operations of distillation we m[illegible] [illegible] the same [illegible] conversion [illegible]

[illegible]

A certain weight of steam at 212° F., condensed at 32° F., gives out	180° of sensible heat. 950° of latent heat.
Amounting together to . . .	. 1130°
The same weight of steam at 250°, condensed at 32°, gives out	218° of sensible heat.
But only	. 912° of latent heat.
Still amounting together to . .	. 1130°
The same weight of steam at 100°, condensed at 32°, gives out	68° of sensible heat.
But now as much as . . .	. 1062° of latent heat.
Making together, as before . .	. 1130°'' [1]

268. If liquids required for their conversion into vapour the me total quantity of heat, there would be no economy in istilling at one temperature rather than another.

269. Regnault has submitted the subject to a rigorous exmination, and his results show that the total amount of heat all temperatures is not the same, but increases by a constant ifference equal to 0·305 for each degree C. He found that e latent heat of steam at 0°C. is 606·5 ; so that the formula r calculating the total quantity of heat in steam at different mperatures becomes—

$$\lambda=606{\cdot}5+0{\cdot}305t,$$

which λ represents the sum of the latent and sensible heat, hile 606·5 is the latent heat of the vapour at 0°, and t the ven temperature.

270. The elastic force of a vapour varies with the conditions nder which the vapour is formed ; if the vapour is produced elow the boiling-point of the liquid, its elastic force is not equal the pressure of the atmosphere ; if the vapour is produced the boiling-point of the liquid, its elastic force is equal the pressure of the atmosphere ; if the vapour is produced bove the normal boiling-point of the liquid, its elastic force xceeds the normal pressure of the atmosphere. Steam of reater tension than the atmospheric pressure is called *high ressure* steam. The tension of steam increases very rapidly ; the temperature of 250·52° F. it is equal to two atmospheres

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[1] Miller's *Chemical Physics*, 2nd edition.

Temp. Fah.	Force. Inches of Mercury.	Temp. Fah.	Force. Inches of Mercury.	Temp. Fah.	Force. Inches of Mercury.
42	0·270	70	0·723	98	1·764
43	0·280	71	0·748	99	1·819
44	0·291	72	0·773	100	1·874
45	0·302	73	0·799	101	1·931
46	0·313	74	0·826	102	1·990
47	0·324	75	0·854	103	2·050
48	0·336	76	0·882	104	2·112
49	0·349	77	0·911	105	2·176
50	0·361	78	0·942	106	2·241
51	0·375	79	0·973	107	2·307
52	0·389	80	1·005	108	2·376
53	0·402	81	1·036	109	2·447
54	0·417	82	1·072	110	2·519
55	0·432	83	1·106	111	2·593
56	0·447	84	1·142	112	2·669
57	0·463	85	1·179	113	2·747
58	0·480	86	1·217	114	2·826
59	0·497	87	1·256	115	2·908
60	0·514	88	1·296	116	2·992
61	0·532	89	1·337	117	3·078
62	0·551	90	1·380	118	3·166
63	0·570	91	1·423	119	3·257
64	0·590	92	1·468	120	3·349
65	0·611	93	1·514	121	3·444
66	0·632	94	1·562	122	3·542
67	0·654	95	1·610	123	3·641
68	0·676	96	1·660	124	3·743
69	0·699	97	1·712		

TABLE II.—*Elastic Force of Aqueous Vapour in Millimetres of Mercury for Centigrade Thermometer.*

Temp. Cent.	Force of the Aqueous Vapour in Millimetres.	Temp. Cent.	Force of the Aqueous Vapour in Millimetres.	Temp. Cent.	Force of the Aqueous Vapour in Millimetres.
0°	4·525	14°	11·882	28°	28·148
1	4·867	15	12·677	29	29·832
2	5·231	16	13·519	30	31·602
3	5·619	17	14·409	31	33·464
4	6·032	18	15·351	32	35·419
5	6·471	19	16·345	33	37·473
6	6·939	20	17·396	34	39·630
7	7·436	21	18·505	35	41·893
8	7·964	22	19·675	36	44·268
9	8·525	23	20·909	37	46·758
10	9·126	24	22·211	38	49 368
11	9·751	25	23·582	39	52·103
12	10·421	26	25·026	40	54·969
13	11·130	27	26·547		

paratus. Two barometer-tubes, A and B (fig. 43), are put into the same basin of mercury, which is placed upon a furnace. The barometer, B, is completely freed from air and moisture; in other words, it is a perfect barometer; the barometer, A, contains a small amount of water above its mercury column. These two barometers are enclosed in a tall glass cylinder filled with water, and a thermometer, T, dips into the water in the centre of the cylinder, which gives the temperature of the liquid. In heating gradually the basin containing the mercury, and consequently the water in the glass cylinder, the water in the tube continues to vaporise, and as the tension of the aqueous vapour augments, the mercury in A falls lower and lower. The depression which is produced in A, below the level in B, for each degree the temperature is increased, is noted upon the scale, E. The apparatus of Dalton can be used so long as the elastic force of the vapour does not exceed the pressure of the atmosphere. When the tension is equal to the atmospheric pressure, the surface of the mercury column will be depressed to the level of the mercury in the basin, and the experiment is at an end.

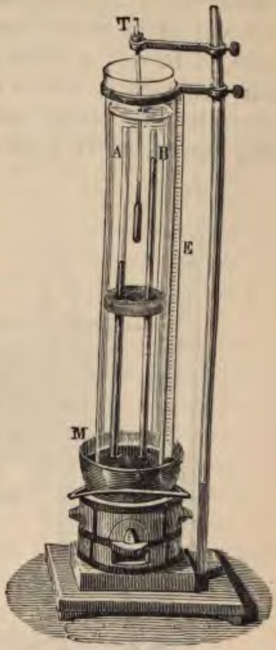

FIG. 43.

275. Regnault employed a form of apparatus which has the advantage of indicating all pressures and temperatures, whether above or below the boiling-point. 'His process consists in boiling water in a vessel under a known pressure, and ascertaining the temperature at which it boils. This method depends upon the principle that when the water boils, the steam it produces will have a pressure precisely equal to that to which the water itself is submitted.

276. 'The apparatus consists of a copper boiler (fig. 44), closed so as to be steam-tight, filled to about a third of its atmosphere at the time supports a column of mercury 30 inches in length, its pressure is equal to 15 lbs. on every square inch; if the vapour of water depresses the mercurial column one inch, its pressure must be equal to $\frac{1}{30}$ of 15 lbs.; if the alcohol depresses the mercury 2 inches, and the ether depresses it 20, the pressure they exercise is found in the same way. It has been proved that the elasticity of the vapour emitted from each liquid increases as the temperature rises, until, at the boiling-point of each liquid, the elasticity of the vapour is equal to the pressure of the atmosphere.[1]

282. Dalton considered that all vapours have the same elasticity at an equal number of degrees above or below their boiling-points; other experimenters have not found this to be perfectly correct, although for short distances above or below the boiling-point it is very nearly true, except in the case of mercury.

283. The quantity of vapour that can be formed from any liquid depends upon the space and the temperature; the same quantity of vapour will pass into the space if it be filled with gaseous matter or with other vapours, or if it be perfectly empty; the only difference being that it requires a longer time to pass into the space which contains other gaseous or vaporous bodies than it would do if the space were empty. 'Different gases and vapours offer no resistance to each other's elasticity; thus, the particles of watery vapour in the air are not subjected to the pressure of the atmosphere, but only influenced by the pressure of the particles of the same kind; and hence, at 32°, when the elasticity of the vapour is only 0·200 inch, it retains perfectly its elastic constitution, though diffused through an atmosphere the elasticity of which may equal thirty inches.'

284. The mixture of a gas and vapour obeys the two following laws; these laws were discovered by Dalton, and are known under the name of the 'Laws of Dalton':—

1. *The amount of vapour which will saturate a given space at a given temperature, and its tension, are the same, whether the space be completely empty or filled with gas.*

2. *The elastic force of a mixture of gas and vapour is equal to the sum of the tensions which each would have separately.*

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[1] E is a graduated scale for measuring the depressions.

285. Some of the aqueous vapour in the air is deposited, we ave before noticed, in the form of dew, rain, hail, or snow when he temperature becomes lowered;[1] and everyone must have bserved, that when a vessel containing cold water is brought nto a warm room, moisture becomes deposited on its outer surface owing to the air surrounding the vessel becoming cooled by the water. The temperature at which the deposition takes place is called the *dew-point*; instruments which are employed to measure the amount or degree of moisture in the air, are termed *hygrometers.* The degree of moisture does not depend on the absolute quantity of aqueous vapour in the air, but on the greater or less distance of the air from its point of saturation. When the air is cold, it may be moist with very little vapour, and, on the contrary, when it is warm, very dry, even with a large quantity of vapour; but we must refer the student for further information on this subject to works on Natural Philosophy.

286. We will close this chapter with instructions for *correcting the volume of gases for the tension of aqueous vapour,* and with some exercises on the subject; but it is necessary first to describe the *apparatus* employed to hold gases for analysis: and the *correction of gases for pressure.*

287. In gas analysis, the gases are *collected* and *measured* in *graduated* tubes over mercury; in Bunsen's method of gas analysis, two gas tubes are employed, one is called the absorption tube (fig. 46), and the other is called the eudiometer (fig. 47). In this latter tube the combustible gases are exploded; there is therefore fused into the closed end of the tube, as shown in the figure, two platinum wires for the purpose of conducting the electric spark. The tubes are first filled with mercury, the mouth of the tube is then closed air-tight with the thumb and inverted, mouth downwards, in the trough containing mercury. The gaseous substance to be analysed is then conveyed into the tube, and it displaces, of course, an equal volume of mercury in the tube. At each stage of the analysis, four observations have to be made: 1st, the height of the mercury in the gas tube; 2nd, the height of mercury in the trough as measured on the scale of the gas tube; 3rd, the temperature; and 4th, the atmospheric pressure.

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[1] The condensation of some of the aqueous vapour in the atmosphere in the form of dew takes place most copiously, as no doubt all my readers will have noticed, during a calm, clear night succeeding a hot day. This is due to the earth's surface substances, especially the vegetation, radiating heat, which is not re-radiated except on cloudy nights, the temperature on the earth's surface is therefore lowered, and consequently some of the aqueous vapour in the stratum of air in contact with the surface is deposited on account of this reduction in temperature; just as moisture is deposited on a glass containing cold water when brought into a warm room.

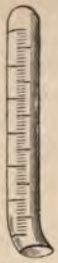
FIG. 46.

FIG. 47.

288. At the moment a gas is weighed or measured, it is necessary to observe the height of the barometer, as this gives the pressure to which the gas is then subjected, provided the liquid over which the gas is collected stands on the *same level* both *in* and *outside* the vessel containing the gas. If, however, the liquid stands *higher* inside than outside the gas-vessel, the gas will be subjected to a pressure less than that of the atmosphere at the time, by the amount necessary to support the column of liquid above its outer level. If the liquid stands *lower* in the gas vessel than outside, the gas is under greater pressure than that of the atmosphere by the amount of pressure which a column of the liquid, equal to the difference in its height, without and within the gas-vessel exercises. In this latter case, the perfect level of the liquid within and without the vessel may readily be restored by raising the gas-vessel. If the fluid stands higher within than outside the vessel, which is generally the case, it is rarely possible in practice to bring it to the same level. A correction has therefore to be made for this inequality of level, which is accomplished in the following way:—Suppose the gas to have been collected over mercury; in order to allow for the dilatation occasioned by the inequality of level, the difference of the two heights must be accurately measured, and the measurement so obtained must be subtracted from the height of the mercurial column in the barometer at the time. A similar correction is required if the gas be standing over water, but it is smaller in amount,—a column of water, 13·6 in height, being equivalent to one inch of mercury. When this correction has been made, a simple calculation enables one to discover the volume any gas would measure at the standard pressure.

Example.—What volume would a gas, standing over mercury, measure at the standard pressure, which measures 60 cubic inches, the barometer at the time standing at 28·75, the level of the mercury *in* the gas-vessel being 1·5 inch higher than that *outside?* By Mariotte's law, the bulk of a gas is inversely as the pressure to which it is subjected. Therefore,—

Standard pressure		Observed pressure		Observed volume		True volume
30	:	28·75 − 1·50 = 27·25	::	60	:	54·5

Exercises on the Correction of Gases for Pressure.

110. What volume would a gas, standing over mercury, measure at the standard pressure, which measures 19 cubic inches, the barometer at the time standing at 29·05, the level of the mercury in the gas-vessel being 0·5 inch higher than that outside?

111. What volume would a gas, standing over mercury, measure at the standard pressure, which measures 20 cubic inches, the barometer at the time standing at 28·05, the level of the mercury in the gas-vessel being 1·0 inch higher than that outside?

112. What volume would a gas, standing over mercury, measure at the standard pressure, which measures 15 cubic inches, the barometer at the time standing at 31, the level of the mercury in the gas-vessel being 0·2 inch higher than that outside?

289. *Correction of gases for the tension of aqueous vapour.*—We learn from Table I., that the elastic force of the vapour of water at 80° F. would depress the mercurial column one inch; or its tension is one-thirtieth of the usual tension of the air; therefore if a gas perfectly dry were saturated with aqueous vapour at 80° F., the vapour would increase the tension of the gas, if it were confined, by one-thirtieth; or if the gas were allowed to expand, the vapour would increase its bulk by one-thirtieth. A gas containing aqueous vapour must therefore occupy more space than if it were perfectly dry, by reason of the pressure which the vapour exercises on the confining fluid. It is, therefore, necessary to take care, in measuring gases, that they are either perfectly dry, or thoroughly saturated with aqueous vapour. To find the actual pressure upon a gas saturated with vapour, we must subtract from the apparent pressure the amount neutralised by the tension of aqueous vapour. Example:—Suppose a gas saturated with aqueous vapour measured 100 cubic inches, the temperature being 50° F., and the atmospheric pressure at the time 29·361 inches; how much space would the gas occupy in the dry state, and at 30 inches pressure? The tension of aqueous vapour at 50° F. is 0·361 (Table I.), therefore the gas is not under the apparent pressure of 29·361, but under the actual pressure of 29·361 − 0·361 = 29·0 inches; therefore,—

Standard pressure		Observed pressure		Observed volume		True volume
30	:	29·361 − 0·361 = 29·0	::	100	:	96·6

290. We have to attend to the following points in the measuring of gases:—1. We must determine as accurately as possible the height of the water or mercury in the vessel containing the gas, above that outside; and the difference must be subtracted from the height of the barometric column at the time. 2. If moist, we must correct for the tension of aqueous vapour. 3. These two corrections being made, and having observed the height of the barometer, we have to subtract from the apparent pressure, the height of the liquid in the gas-vessel and the tension of the vapour, to get the actual pressure to which the gas is subjected at the moment it is measured. 4. We must correct the gas for temperature.[1] As the volumes of gases can be compared only if measured at the same temperature, under the same pressure, and in the same hygroscopic state; the temperature is generally reduced to 60° F.,[2] the hygroscopic state of the gases to 0, and the pressure to 30 inches.[3]

291. If it is desired to find the weight of a gas which has been measured, it is necessary to ascertain the weight corresponding to the volume found, before the proportional amount

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[1] It does not matter in what order the corrections are made.

[2] This, as we have before stated, is the English standard; the continental one is 0° C. = 32° F.

[3] The standard pressure adopted on the Continent is 760 millimetres = 29·922 inches.

139. What are the fixed points from which a thermometer is graduated? State why they have been selected. What degree on the Fahrenheit scale corresponds with $-7°$ Cent.; and what degree on the Cent. scale with 70° F.?

140. What occasions the *singing* noticed in liquids before they begin to boil?

141. How do you account for sea-fogs?

142. Why do some liquids on exposure to the air increase in volume?

143. If a closed vessel filled with steam at 212° F. at the ordinary atmospheric pressure be raised to 300°, will the presence or absence of water, as well as steam, within the vessel affect the expansive force?

144. Give the true definition of the boiling-point of a liquid.

145. Define what is meant by specific heat; and give the specific heat of three or four familiar substances. *Exercise.*—Given two vessels of equal size, one filled with water at 10° C.; the other with mercury at 10° C.; find out how much heat each liquid must receive before it begins to boil, the weight of the water being 2 lbs., the specific heat of mercury ·033, and its boiling-point 350° C.

146. Why does a standing leap fall short of a running one?

147. If I wished to evaporate a liquid spontaneously, whether would it be more advisable, in order that it should evaporate as speedily as possible, to place it in a vessel which causes the depth to be greater than the surface; or in a vessel in which the surface of the liquid is greater than its depth?

148. How is deposition of dew affected by the presence or absence of clouds in the sky?

149. What is the theory of fractional distillation? How is the process carried out?[1]

CHAPTER V.

ABSOLUTE DENSITY — SPECIFIC GRAVITY — DIFFERENCE BETWEEN DENSITY AND SPECIFIC GRAVITY—SPECIFIC GRAVITY OF LIQUIDS —EXERCISES—SPECIFIC GRAVITY OF SOLIDS—EXERCISES—TABLES OF SPECIFIC GRAVITIES, THEIR USE—EXERCISES—THE HYDROMETER, ITS USE—SPECIFIC GRAVITY OF MIXTURES—EXERCISES—EXPERIMENTAL EXERCISES—WEIGHT AND SPECIFIC GRAVITY OF GASES—EXERCISES—SPECIFIC GRAVITY OF VAPOURS—EXERCISES—QUESTIONS.

292. IN Chapter III. the meaning of the word MASS was explained; it meant, it was stated, the quantity of matter in a substance, and that it was estimated in the terms of some standard. We have now another term, ABSOLUTE DENSITY, to explain before we can enter satisfactorily upon the subject-matter of the present chapter. We have constantly, in every-day life, to notice and speak of one substance being *lighter* than another. Now we do not mean, by that expression, that one ounce, or one pound, of the one substance weighs less than an ounce or a pound of the other substance ; but we mean, that the lighter of the two substances we are alluding to, weighs less for an *equal volume* or *bulk* than the other substance ; one cubic inch of cork, for example, weighs less than a cubic inch of iron. We say, in like manner, that, of the two liquids—water and mercury—water is the lighter of the two, because we know by experience that, if we were to weigh equal volumes of the two liquids, mercury would prove the heavier of the two bodies ; then, that is the *more dense* which in *equal volume* contains the *greater mass*. Different bodies have, therefore, as we daily notice, different densities ; and the simplest way of determining the *density* of a substance is to find the mass of a known volume of it by weighing it, and to divide the mass obtained by the volume ; we thus obtain the quantity of matter in unit-volume of any substance, and this is defined as the *absolute density* of that substance. If, then, we denote the mass by M, the volume by V, and the density of the body by D, we have

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[1] The student must now return to Chapter II. and complete the study of it, commencing at par. 144.

$$D = \frac{M}{V}$$

293. Having explained what is meant by the term *absolute density*, we have now to explain what is meant by the *relative density*, or, as it is generally termed, the *specific gravity*, of a substance ; these two terms are applied to the ratio between the masses of equal volumes of a given body and of some standard substance ; and this forms the subject-matter of the present chapter.

294. Density means, then, comparative *mass*, and specific gravity comparative *weight*. These expressions, although really relating to distinct things, are often used indifferently in chemical writings, and without practical inconvenience, since mass and weight are directly proportional to each other.

295. To arrive at the relative weights of equal volumes of different substances, it is necessary to fix upon one as the standard of comparison : distilled water has been selected as the standard for solid and liquid bodies, its density being reckoned as 1000 or 1. The specific gravity of substances heavier than water is consequently represented by a higher number, and those which are lighter by a lower number than 1000 or 1.

296. The specific gravity of the metal platinum is stated to be 21·5; by this is meant that for *equal volumes* platinum is 21·5 times heavier than water. Thus, one cubic foot of water, at the temperature of 60° F., weighs 1,000 ounces; therefore as platinum, bulk for bulk, is 21·5 times heavier than water, one cubic foot of platinum, at 60° F., will weigh 21,500 ounces. Again, comparing equal volumes, turpentine is lighter than water in the proportion of 792 to 1,000; therefore, as one cubic foot of water, at the temperature of 60° F., weighs 1,000 ounces, a cubic foot of turpentine at that temperature will weigh 792 ounces.

297. In order to calculate the specific gravity of a body it is sufficient to determine its weight and the weight of an *equal volume* of water, and then *divide the weight of the body by the weight of water; the quotient is the specific gravity of the body*, as compared with that of water taken as unity. The formula for specific gravity is, therefore,

$$S = \frac{W}{w}$$

W representing the weight of the body, w the weight of an equal volume of water, and S the specific gravity of the substance.

298. It is desirable at the outset to inform the student that the experiment is generally made in this country at a temperature of 60° F., and the standard barometric pressure for gases and vapours is 30 inches; in almost all other countries the French standards of pressure and temperature are employed; viz., 760 millimetres, or 29·922 inches for the pressure; and 4° C., or 39°·2 F., for temperature, for at that temperature water *attains*, as we have already learned, the point of maximum density.

299. Bottles (fig. 48) are made for the express purpose of determining the sp. gr. of liquids; they are made to hold 1,000, 500, 250, &c., grains or grammes of distilled water. A weight is usually sold with them, which counterpoises the bottle full of distilled water, and upon which is marked the quantity of distilled water the bottle holds. All that is required in estimating the sp. gr. of any other fluid is to fill the bottle with it completely ; to close the bottle with the stopper in such a way that no bubbles of air are left in it ; then, after wiping it perfectly dry, to ascertain its weight, by observing how much heavier or lighter it is than the weight of water the bottle is capable of holding, and adding or subtracting accordingly; and the result, divided by the weight of water the bottle is capable of holding, will be the sp. gr.

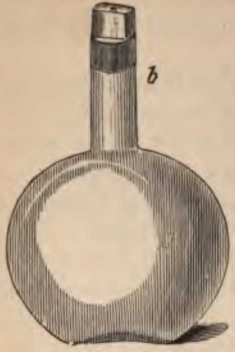

Fig. 48.

300. A bottle capable of holding 500 grains of distilled water will hold 922·5 grains of sulphuric acid; the latter weight divided by the former will give the sp. gr. of the acid, water being taken as 1 ; thus, $\frac{922{\cdot}5}{500}=1{\cdot}845$ sp. gr. of the acid : again, the same bottle, or one of like size, will hold only 396 grains of turpentine, therefore $\frac{396}{500}=0{\cdot}792$ sp. gr. of the turpentine.

301. As the volume of liquids expands with increase of temperature, the quantity of ponderable matter in a given volume becomes thereby proportionably diminished ; it is therefore requisite, in determining the sp. gr. of substances, to conduct all the experiments at one uniform standard of temperature. For the same reason, the bottle should never come in contact with the naked hand during the experiment ; it should be protected by a cloth during the wiping and drying of it and its transfer to the balance.

Exercises on the Calculation of the Specific Gravity of Liquids.

150. If a bottle, capable of holding 300 grains of distilled water, holds 260 of some other liquid, what is the sp. gr. of that liquid ?

151. If a bottle, capable of holding 360 grains of distilled water, holds 320 grains of some other liquid, what is the sp. gr. of that liquid ?

152. If a bottle, capable of holding 400 grains of distilled water, holds 470 grains of some other liquid, what is the sp. gr. of that liquid ?

153. If a bottle, capable of holding 700 grains of distilled water, holds 600 grains of some other liquid, what is the sp. gr. of that liquid ?

302. *The weight of any substance may be calculated if the volume and specific gravity be known.* The weight of the volume-unit, be it a gallon, a cubic inch, a cubic foot, &c., as the case may be, must first be determined, and this is arrived at by *multiplying the weight of the same volume-unit of water by the specific gravity of the substance. The product is the weight of the volume-unit of the substance.* Thus, if we represent the weight of the volume-unit of water by w, and the specific gravity of the substance by S, and the weight of the volume-unit of the substance by W; then

$$W = S \times w.$$

Example: 1. Required the weight in pounds of twenty gallons of sulphuric acid, the sp. gr. of which is 1·85.

Now, the weight of a gallon of water, which is in this case the volume-unit, is 70,000 grains, or ten pounds avoirdupois; therefore—

10 lbs. × 1·85 = 18·5 lbs. weight of 1 gallon of sulphuric acid.

18·5 lbs. × 20 = 370·0 lbs. weight of 20 gallons.

2. Required the weight in pounds of a dry block of fir containing 50 cubic feet, the sp. gr. of the wood being 0·47.

Now the weight of a cubic foot of water, which is in this case the volume-unit, is 62·5 lbs.; therefore,

62·5 lbs. × 0·47 = 29,375 lbs. weight of 1 cubic foot of fir.

29,375 lbs. × 50 = 1468·75 lbs. weight of 50 cubic feet of fir.

Exercises on the Calculation of the Weight of Liquids from their Volume and Specific Gravity.

154. Required the weight in pounds of a gallon of linseed oil, its sp. gr. being 0·953.

155. Required the weight in pounds of a gallon of turpentine, its sp. gr. being 0·792.

156. Required the weight of a gallon of the water of the Dead Sea, its sp. gr. being 1·172.

157. Required the weight of a gallon of vinegar, its sp. gr. being 1·063.

158. Required the weight of a cubic inch of mercury, the sp. gr. of which is 13·59; a cubic inch of water weighing 252·458 grains.

303. When a solid is immersed in water or any other liquid, it displaces a quantity of the fluid equal to its own bulk (as no two bodies can occupy the same space at the same time); the *weight* of the displaced water being determined, as well as that of the solid, the sp. gr. of the solid is the quotient obtained by dividing the latter weight by the former.

304. The specific gravity of solids in the form of powder or small fragments is determined by weighing them in air and then introducing the weighed particles into a sp. gr. bottle filled with water; a quantity of water overflows, equal in bulk to the solid particles introduced; the bottle containing the substance weighs, therefore, less than the bottle filled with water and the substance when they are weighed separately, the difference being the weight of the volume of water expelled by the solid. Thus, if we weigh 200 grains of sand and then introduce it into a sp. gr. bottle holding 500 grains of water, if no water were expelled, the water and sand together would weigh 700 grains; but instead of that it only weighs 624 grains, the difference, 76, being the weight of the water expelled, therefore $\frac{200}{76} = 2{\cdot}63$, the sp. gr. of the sand.

305. Great care must be taken to prevent air being admitted when the solid is added to the water in the bottle, and if any should be introduced every bubble must be expelled before weighing, if we would obtain accurate results.

Exercises on the Calculation of the Specific Gravity of Solids.

159. If 100 grains of a solid be introduced into a bottle holding 500 grains of water, and if after the introduction of the solid the bottle weighs 560 grains, what is the sp. gr. of the solid?

160. If 160 grains of a solid be introduced into a bottle holding 400 grains of water, and if after the introduction the bottle weighs 500 grains, what is the sp. gr. of the solid?

161. If 300 grains of a solid be introduced into a bottle holding 700 grains of water, and if after the introduction the bottle weighs 850 grains, what is the sp. gr. of the solid?

306. When the substance whose sp. gr. has to be determined is soluble in water, it is necessary to employ some other liquid in which it is insoluble; alcohol, naphtha, turpentine, oil, &c., are the liquids generally employed when water cannot be used. The sp. gr. of the solid in reference to the fluid employed is first determined; then the sp. gr. (if it is not known) of the fluid in reference to water; the two numbers thus obtained, qual to that of the fluid, it will remain sus- wherever it may be placed. If the solid y than the liquid, it will sink, because the ht is greater than the upward pressure of olid has a *less* density than the liquid, the f the latter will exceed the weight of the ll therefore be forced up to the surface of rce equal to the difference between the olumes of the solid and the liquid. It is hese effects by stating *that a solid which the surface of a liquid loses as much of its o that of the liquid displaced.*

nciple, which constitutes the foundation of g to immersed and floating bodies, is known e principle of Archimedes.

gr. of solids in the mass is determined by by means of a horsehair nder one of the scale- then weighed in air, the previously taken care to or loosely adhering par- of pure water at 60° F. is a manner that the solid, from the scale-pan, can nearly as possible in the sel (fig. 50), and covered an inch of water (if air- adhere to the solid when he water, they must be re- of a feather); it is now nd the weight will be less ubtract the weight in water from the weight in mainder will be the weight lost on immersion, ords, the weight of an equal volume of water; ght in air by the weight lost on immersion, the the sp. gr. of the solid.

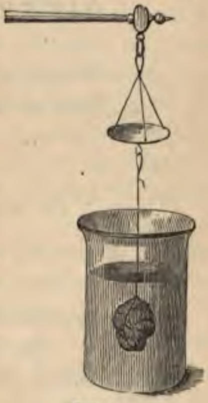

FIG. 50.

e Calculation of the Specific Gravity of Solids.

he sp. gr. of a solid which weighs 36 grains in air and liquid as is equal in volume to the solid immersed ; and if we can collect and weigh the volume of the liquid displaced, and we have weighed the solid previous to its immersion, their weights are of course the *relative* weights of equal volumes of the solid and liquid. We have now to consider another of the effects attending the immersion of solids in liquids. It is this:

308. *The solid will be pressed upwards with a force equal to the weight of the liquid it displaces.* For when a solid is wholly immersed in a liquid, the *lateral* pressures *a* and *b* are equal, and contrary to each other, and therefore *neutralise* each other; but the vertical pressures on the faces *d* and *c* are unequal, the *upward* pressure being greater than the *downward* ; therefore the solid, as a consequence, has a tendency to move upwards. Thus suppose that a cube is immersed in a vessel of water (fig. 49), and that four of its sides are arranged vertically. These four sides, presenting the same surface to the liquid, and being immersed to the same depth, experience *equal* pressures, and as the sides are opposite to each other, two and two, the lateral pressures must be *contrary* as well as *equal*; hence they are neutralised, or in equilibrium. With regard, however, to the pressures which act vertically on the horizontal sides *d* and *c*, it is evident that the downward pressure on *d* is equal to that of the weight of a column of water having the side *d* for its base, and *d n* for its height : and that the upward pressure on *c* is equal to that of the weight of a column of water having the side *c* for its base and *c n* for its height. The cube, therefore, tends to rise under the pressure of a force equal to the difference of these pressures, which is equal, evidently, to the weight of a column of water having the same base and the same height as the cube; consequently, this pressure is equivalent to the weight of water displaced by the body immersed.

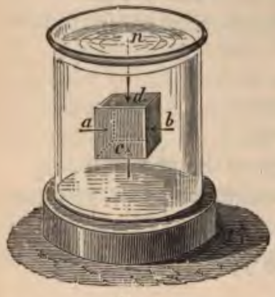

FIG. 49.

309. If the weight of the solid be *equal* to that of the mass of fluid which it has displaced; in other words, if its density or specific gravity be equal to that of the fluid, it will remain suspended in the liquid wherever it may be placed. If the solid has a *greater* density than the liquid, it will sink, because the pressure of its weight is greater than the upward pressure of the liquid. If the solid has a *less* density than the liquid, the upward pressure of the latter will exceed the weight of the solid; the solid will therefore be forced up to the surface of the liquid with a force equal to the difference between the weights of equal volumes of the solid and the liquid. It is usual to express these effects by stating *that a solid which is plunged beneath the surface of a liquid loses as much of its weight as is equal to that of the liquid displaced.*

310. This principle, which constitutes the foundation of the theory relating to immersed and floating bodies, is known by the name of the principle of Archimedes.

311. The sp. gr. of solids in the mass is determined by suspending them by means of a horsehair from the hook under one of the scale-pans; the mass is then weighed in air, the operator having previously taken care to remove any dust or loosely adhering particles. A vessel of pure water at 60° F. is arranged in such a manner that the solid, still suspended from the scale-pan, can be immersed as nearly as possible in the centre of the vessel (fig. 50), and covered by at least half an inch of water (if air-bubbles should adhere to the solid when immersed in the water, they must be removed by means of a feather); it is now to be weighed, and the weight will be less than in air; subtract the weight in water from the weight in air, and the remainder will be the weight lost on immersion, or, in other words, the weight of an equal volume of water; divide the weight in air by the weight lost on immersion, the quotient will be the sp. gr. of the solid.

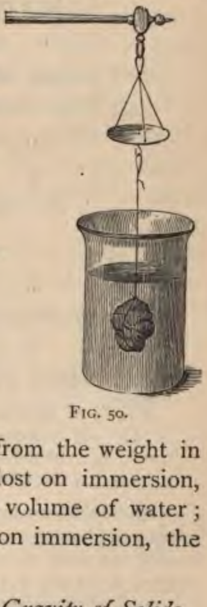

FIG. 50.

Exercises on the Calculation of the Specific Gravity of Solids.

164. Required the sp. gr. of a solid which weighs 36 grains in air and 26 in water.

the weight of the whole mass of the substance *must be divided by the weight of one volume-unit of that substance* ; the method for finding the weight of a volume-unit of a substance has already been given.

Exercises for computing the Weight from the Volume and the Volume from the Weight of Substances.

173. What will be the weight of a block of limestone containing 12 cubic feet; 1 cubic foot of water weighing 62·5 lbs., the sp. gr. of the stone being 2·64?

174. How many cubic feet are there in a block of coal weighing 1 cwt., its sp. gr. being 1·232?

175. The sp. gr. of bar iron is 7·788; required the weight of a cubic foot.

176. The sp. gr. of flint glass is 3·329; required the weight of a cubic foot.

177. The sp. gr. of oak is 0·845; required the weight of a cubic foot.

178. The sp. gr. of cork is 0·240; required the weight of a cubic foot.

179. The sp. gr. of ice is 0·930; required the weight of a cubic foot.

180. The sp. gr. of silver is 10·474; required the weight of a cubic inch.

181. A solid weighs 49 grs. in air, and 42 grs. in water; required the weight of a cubic foot of the substance.

182. A block of marble, of which the sp. gr. is 2·4, weighs 5 cwt.: find its volume, the weight of a cubic foot of water being 62·5 lbs.

315. The sp. gr. of a liquid may also be determined by the following method:—Take a piece of glass, or some other solid heavier than and insoluble in water, determine by the method described at par. 311 the weight of the volume of water at 60° F. the solid displaces; let this weight be represented by w. Then determine by the same method the weight of the volume the solid displaces of the liquid whose sp. gr. we wish to find; let this weight be represented by w'. Divide the loss of weight in the liquid by the loss of weight in water, the quotient will be the sp. gr. of the liquid.

$$s = \frac{w'}{w}$$ [1]

---

[1] The specific gravity of the liquid may also be found by multiplying the difference in weight of the solid in air and in the liquid by the specific gravity of the solid, and dividing the product by the absolute weight of the solid; the result will be the specific gravity of the liquid. For the first of these quantities, namely the difference in weight of the solid in air and in the liquid, is the weight of the volume of the liquid, which is equal to the volume of the solid; and the second of these quantities is the weight of a volume equal to the volume of the solid.

An Indirect Method for determining the Specific Gravity of Substances.

183. A solid whose weight in air is 60 grs. weighs 40 grs. in water and 30 grs. in sulphuric acid; required the sp. gr. of the acid.

184. A glass ball, weighing 10 grs., loses 3·64 grs. in water, and 2·88 grs. in alcohol. What is the sp. gr. of the alcohol?

185. If a piece of metal weigh in air 300 grs. more than in water and 252 grs. more than in spirit, what is the sp. gr. of the spirit?

186. A body weighs 252 grs. in air, 36 in water, and 63 in spirit; what is the sp. gr. of the body and of the spirit?

187. A piece of copper (sp. gr. 8·85) weighs 446·3 grs. in a liquid and 490 grs. out of it; required the sp. gr. of the liquid.

188. A piece of gold (sp. gr. 19·36) weighs 140 grs. in a liquid and 159·7 grs. in air; what is the sp. gr. of the liquid?

316. For commercial purposes the sp. gr. of liquids is determined by means of an instrument called the hydrometer; this instrument is not so accurate as the sp. gr. bottle, but it is preferable to it on account of its simplicity and the rapidity with which the experiment can be made with it, and at the same time it is sufficiently accurate for all commercial purposes.

317. 'The hydrometer (fig. 51) in general use consists of a floating vessel of thin metal or glass, having a weight beneath to maintain it in an upright position, and a stem above bearing a divided scale. The use of the instrument is very simple. The liquid to be tried is put into the glass jar represented in the figure, and the instrument floated in it. It is obvious that the denser the liquid the higher will the hydrometer float, because a smaller displacement of fluid will counterbalance its weight. For the same reason in a liquid of less density, it sinks deeper. The hydrometer comes to rest almost immediately, and then the mark on the stem may be read off. The graduation of the stem is very commonly arbitrary, two or three different scales being unfortunately in use. These may be sometimes reduced, however, to the true numbers expressing

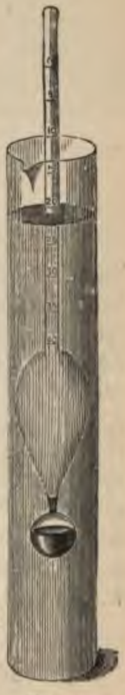

FIG. 51.

the sp. gr. by the aid of tables of comparison drawn up for the purpose.'[1]

318. The sp. gr. of a mixture of two liquids can be found, if the sp. gr. of each of the liquids comprising the mixture be known, and if the proportions *by volume* in which the two liquids have been mixed together be known, and assuming that no alteration in volume occurs on mixing the two liquids:—Multiply the volume of each of the liquids by its sp. gr.; add the two products together, the sum will be the *weight* of the *mixture*; then add the volumes of the liquids together, and the sum will be the *volume* of the *mixture*; then divide the *weight* of the *mixture* by its *volume*, and the quotient will be the *specific gravity* of the mixture. To express this process in an algebraic form, let s' represent the sp. gr. of one of the liquids and v' its volume, and let s'' represent the sp. gr. of the other liquid and v'' its volume, and let S represent the sp. gr. of the mixture, V its volume, and W its weight, then—

$$S = \frac{W}{V} = \frac{v'\,s' + v''\,s''}{v' + v''}$$

319. In like manner, the sp. gr. of a mixture (whether of liquids or solids) can be found, if the sp. gr. of each of the two substances composing the mixture be known, and also the proportion by *weight* in which they are mixed, assuming as before that no contraction takes place; for, find the volume of each of the two substances by dividing its weight by its sp. gr., add the two volumes thus found together, and the sum will be the volume of the mixture; add also the weight of each of the two substances together, and the sum will be the weight of the mixture; divide this weight by the volume as previously found, and the quotient will be the sp. gr. of the mixture. Thus, let w' represent the weight of one of the solids, and w'' the weight of the other solid, then—

$w' + w'' = \mathrm{W}'$, the *weight* of the mixture.

$\frac{w'}{s'} + \frac{w''}{s''} = \frac{\mathrm{W}}{\mathrm{S}} = \mathrm{V}$, the *volume* of the mixture.

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[1] Fownes' *Elements of Chemistry.*

Therefore, dividing the first of these expressions by the second,

$$\frac{w'+w''}{\frac{w'}{s'}+\frac{w''}{s''}}=\frac{\text{W}}{\text{V}},\text{ the specific gravity sought.}$$

320. *To find the weight of each of two substances when combined in one mass.*—The weight of each of two substances when combined in a mass, can be found, if the weight of the mixture be known, and also the specific gravity both of the mixture and of each of its components. Let S represent the specific gravity of the mixture and W its weight, and s' and s'' the specific gravity of the two components : then,

$$w=\text{W}\frac{(\text{S}-s'')s'}{(s'-s'')\text{S}}$$

the weight of one of the substances in the mixture ; and

$$\text{W}-w'=w'',\text{ weight of the other substance.}$$

Exercises for finding the weight of each of two substances when combined in one mass.

189. If the specific gravities of gold and silver are 19·35, 10·51 respectively, find the sp. gr. of a mixture consisting of seven parts by weight of gold, and six of silver.

190. An alloy of gold and silver weighs 10 grs. in air, 9·37 grs. in water ; what are the proportions of gold and silver ?

191. An alloy of copper and silver weighs 37 grs. in the air, and loses 3·66 grs. when weighed in water. What are the proportions of silver and copper? The sp. gr. of copper = 8·85.

192. British standard gold contains 11 parts by weight of pure gold, and 1 part of copper ; required its sp. gr.

193. A mixed mass of gold and quartz (the sp. gr. of the quartz being 2·6) weighs 20 lbs., its sp. gr. is 8·5 ; how much gold does it contain ?

194. A fluid of sp. gr. 1·12 is mixed with three times its bulk of another fluid having sp. gr. of 1·7. Supposing the fluids do not act chemically on one another, and there is no contraction, what will be the sp. gr. of the mixture ?

321. We have now to describe the methods pursued in determining the sp. gr. of gases ; but before doing so we will gave a few easy experimental exercises which the student will find no difficulty in performing.

Experimental Exercises.

195. Determine the sp. gr. of a sample of olive oil, by means of the sp. gr. bottle.

196. Determine the sp. gr. of a sample of hydrochloric acid by means of the hydrometer.

197. Determine the sp. gr. of marble both in powder and in lump.

198. Determine the sp. gr. of bees'-wax.

322. The determination of the sp. gr. of gases is very simple in theory, but is an exceedingly delicate operation to perform. Air at 60° F., and the barometer standing at 30 inches, is the English standard of comparison; but the standard of temperature and pressure adopted on the Continent and in England for scientific purposes, is 0° C. for the temperature, and 760 millimetres for the pressure.

323. Before describing the operation itself, we will describe

(1) *How to find the weight of any gas or vapour, its sp. gr. being known.* To accomplish this we have simply to multiply the weight of an equal volume of air by the specific weight of the gas or vapour; the product will be the weight of the volume, at the standard temperature and pressure of the gas or vapour.

Example: What is the weight of 100 cubic inches of hydrogen, its sp. gr. being ·0694?

$30{\cdot}935 \times {\cdot}0694 = 2{\cdot}147$ grains weight of 100 cubic inches of H.

(2) *How to find the volume of a given weight of any gas or vapour.* The weight of some volume (say a cubic inch) of the gas or vapour must first be ascertained by the preceding rule; then the given weight must be divided by the weight of the cubic inch; the quotient will be the volume in cubic inches of the given weight of the gas, at the standard temperature and pressure.

Example: What is the volume of 2·147 grains of hydrogen, its sp. gr. being 0·0694?

$$\frac{2{\cdot}147}{{\cdot}02147} = 100 \text{ cubic inches.}$$

Exercises for finding the weight of a given volume of a gas or vapour from its specific gravity.

199. What is the weight of a cubic inch of oxygen, its sp. gr. being 1·1057?

200. What is the weight of a cubic inch of nitrogen, its sp. gr. being 0·9713?

201. What is the weight of a cubic inch of carbon dioxide, its sp. gr. being 1·529?

202. What is the weight of a cubic inch of gaseous ammonia (NH_3), its sp. gr. being 0·59?

203. What is the volume of 54 grains of chlorine, its sp. gr. being 2·44?

204. What is the volume of 45 grains of carbon monoxide, its sp. gr. being 0·967?

324. It is necessary, in order that the student may fully understand Regnault's method for the determination of the density of gases, to direct his attention in the first instance to the *buoyancy of the air*. The general fact, that air, like liquids, buoys up all bodies immersed in it, may be illustrated by means of the apparatus represented in fig. 52. It consists of a closed globe, suspended to one arm of a delicate balance, equipoised by a weight suspended to the other. The two are in equilibrium in the air, but only because the globe, being larger than the weight, is buoyed up by a greater force. If, now, the apparatus is placed upon the plate of an air-pump and covered with a glass bell, we shall find, on removing the air, that the globe will preponderate, as is shown in the figure. By removing the air we increase the apparent weight both of the globe and of the counterpoise by just the weight of the air displaced by each; but as the globe is much the largest, we increase its weight more than that of the smaller brass counterpoise, and hence the result. If we allow the air to re-enter the bell, it will buoy up the globe, as before, so much more than the counterpoise as to restore the equilibrium.

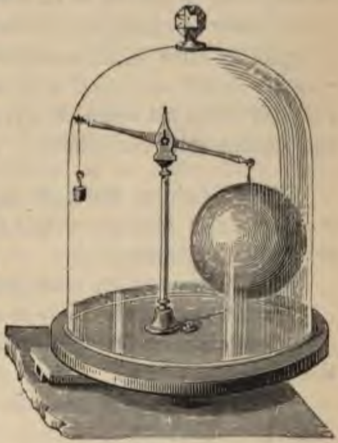
Fig. 52.

325. An important consequence of the principle just illustrated is evident. The balance does not give us the true relative weight, W, of a body, but a slightly different weight depending on the weight of air displaced by the body, compared with the weight of air displaced by the brass or platinum weights used in weighing. As the volume of these weights is generally less than that of the body, the weight indicated by the balance is almost always too small; but when the volume of the weights is greater than that of the body, the weight indicated by the balance is too large. When the two volumes are equal, the balance will indicate the same weight in air as in a vacuum. It is easy to ascertain the correction which it is necessary to add or to subtract from the weight of a body in air in order to obtain its true weight.[1]

326. It must be remembered that the brass and platinum weights which are used in delicate determinations of weight are only standard when in a vacuum. Let us, then, represent the various values as follows:—

W^1=Weight of the body in air as estimated by standard weights; and also the weight of the standard weights themselves in a vacuum.

V^1=Volume of the standard weights in cubic centimetres.

V=Volume of the body in cubic centimetres.

w=Weight of one cubic centimetre of air at the time of the weighing.

W=Weight of the body in a vacuum—which we wish to find.

We can now easily deduce the following values:—

V^1w = Buoyancy of air on the weights.

Vw = Buoyancy of air on the body.

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[1] 'The weight of the displaced portion of air may be easily ascertained if the specific gravity of the body be known; for, from the observed weight of the body, we can calculate directly the weight of an equal bulk of water, and $\frac{1}{817}$th of this weight will give the weight of a corresponding bulk of air at mean temperature and pressure. This weight must be added to that actually found; at the same time, a similar and opposite correction will be required for the metallic weights used in the experiment, because they will also appear to be lighter than they really are; and an amount of weight greater than the true one will be required to effect the counterpoise. If, therefore, the weights have the same specific gravity as the body counterpoised, the two corrections will neutralise each other; but if, as in weighing gases, there is a great difference between them, the correction will be one of importance. The true weight sought will be thus obtained:—*Add to the weight of the body in air, the weight of the bulk of air which it has displaced, and deduct from this the weight of the bulk of air displaced by the weights employed.*'—Miller.

$W^1 - V^1w =$ Actual weight of standard weights in air.
$W - Vw =$ Actual weight of body in air.

Since these weights just balanced each other, we have—

$$W - Vw = W^1 - V^1w\ ;\ W = W^1 + w\,(V - V^1).$$

327. The correction $w\,(V-V^1)$ which must be made to the weight determined by a balance in air in order to obtain the weight in a vacuum, is evidently additive when the volume of the body is greater than that of the weights, and subtractive when these conditions are reversed. When the volumes are equal, the correction becomes 0.

328. In all ordinary cases of weighing, the correction is so small that it may be neglected without sensible error; but it becomes of the greatest importance in determining the weight of a gas. In such cases, we have to determine the weight of a large glass globe, when completely vacuous, and when filled with gas; and it not unfrequently happens that the buoyancy of the air is greater than the weight of the gas itself, and it is always a considerable part of it. If the buoyancy of the air is the same when the globe is weighed in its vacuous condition and when filled with gas, it would not affect the weight of the gas, which would be obtained by subtracting the first weight from the last. But, unfortunately, the buoyancy is constantly changing; and it is, therefore, necessary to determine the amount carefully at each weighing, and reduce the weights of the globe in the two conditions to what they would be if the experiments had been made in a vacuum.

329. When the temperature is 0° C., and the barometer stands at 760 mm. and when the air contains neither vapour of water nor carbon dioxide, w is equal to 0·001293 grams. Were the atmosphere always in this condition, nothing would be easier than to calculate the actual weight of a body from the weight found by weighing in this normal atmosphere. But this is far from being the case; for the temperature, the pressure, and the composition of the atmosphere are changing at each moment, and the value of w varies with all these atmospheric changes.

330. It is frequently possible to conduct the process of weighing in such a way that the correction for the buoyancy of the atmosphere, always somewhat uncertain, may be avoided. For example, in weighing a gas, instead of equipoising the glass globe, when empty, by means of ordinary weights, we may equipoise it by means of a second globe, hermetically closed and having the same volume as the first, in the manner represented in fig. 53. It is evident that in this case, whatever may be the buoyancy of the atmosphere, it will equally affect both globes, and we shall only have to consider the buoyancy of the air on the small weights necessary to restore the equilibrium after the globe is filled with the gas to be weighed; but this is so small that it may always be neglected.[1]

332. The most accurate method of determining the specific gravity of a gas is due to Regnault; it consists in determining with the apparatus, fig. 53, the weight of the given gas which a large balloon will contain at 0° C. and 760mm. and then dividing this weight by that of an equal volume of air previously determined in the same way. The balloon in which the gas is weighed is counterpoised, not entirely by weights, but chiefly, as we have already stated, by another balloon of the same kind of glass and of the same bulk. To do this, it is necessary to determine the volume of air displaced by the balloon in which the gas is to be weighed, and which we will call A. It is to be filled with water, and then weighed in water of the same temperature. The apparent weight of the balloon in water is so small that this may be effected on the balance to be used for the densities. The balloon is now to be removed, and weighed, after wiping, but still filled with water, on a strong balance capable of indi-

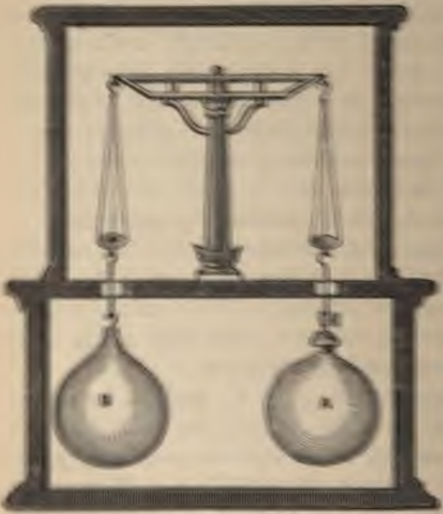

Fig. 53.

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[1] Cooke's Chemical Physics.

cating one decigramme. The difference in the two weighings is equal to the weight of water displaced by the external volume of the balloon. Another balloon, B, fig. 53, of nearly the same capacity as A, but, preferably, a little smaller than that of the first and its stopcock, and made of the same kind of glass, is taken, and has connected to it a brass mounting, terminating in a hook, to enable it to be suspended from the balance-pan. If the united weight of the water displaced by the balloon, B, and its mounting, is less by *n* grammes than the weight of the water displaced by the balloon A, we must attach to B a glass tube closed at each end, and having an exterior bulk = *n* cent. cub. of water.[1]

332. The engraving, fig. 53, shows the method of suspending the balloons beneath the scale-pans of the balance, in a chamber closed with glass doors to prevent currents of air. The balance, although large, is of great delicacy; for when charged with one kilogramme in each pan, it permits, with certainty, an appreciation of half a milligramme. The balloons employed have a capacity of about 10 litres.

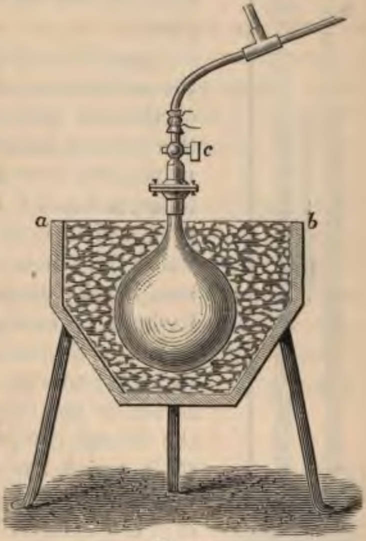

FIG. 54.

333. The balloon A has a stopcock attached to it, which allows it to be connected either with a three-way tube communicating with the gas-holder, or with the air-pump. The air having been removed as completely as possible, the gas is allowed to enter; but, as a small amount of air still remains, the operation is to be repeated twice. Previous to the third time of filling the globe, it is, after as complete exhaustion as possible, to be placed in a case, *a b*, fig. 54, and covered with melting ice; the cock being opened, the globe is allowed to fill with gas, and when full, a momentary communication is made with the atmosphere to equalise the pressure; the cock, *c*, is closed, the globe removed, wiped with a damp cloth to prevent electrical excitation, which might cause serious errors in the weighing, and suspended on the balance. It is not weighed until two hours have elapsed, so as to permit the temperature to become the same as that of the balance-case, and thus obviate currents of air, and also that its surface may be covered with the normal amount of humidity. After careful weighing the balloon is placed anew in the case, *a b*, fig. 54, surrounded with ice, and the gas removed by the pump.

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[1] Because 1 cent. cub. of water weighs 1 gramme.

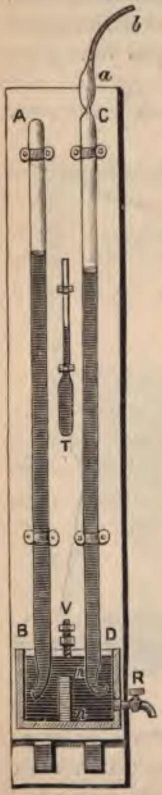

FIG. 55.

334. It is necessary now to ascertain both the atmospheric pressure and the elastic force of the gas remaining in the balloon. For this purpose, an instrument called a barometric manometer is made use of. It consists of two tubes, A B and C D, fig. 55, attached to a support, which is secured perpendicularly to a wall. The tube, A B, is a barometer of 20 mm. interior diameter; the metal in the tube having been carefully boiled, it is inverted in a cistern of dry mercury. This cistern is a box divided into two parts, the smaller of which serves for the cistern of the barometer. Into the second compartment is plunged the tube, C D, which has the same diameter as A B. C D is capable, by means of a leaden tube, *a b*, of being placed in communication with the balloon.

335. When it is intended to ascertain the atmospheric pressure by means of this instrument, mercury is poured into the cistern until it rises above the level of the division, *m n*. The double-pointed screw V is then adjusted, until its lower end just touches the surface of the mercury. If now we measure, y means of a cathetometer, the difference of level between the rface of the mercury in the barometer and the upper point f the screw, and add to this the length of the screw previously certained, we have the height of a column of mercury exactly lancing the atmospheric pressure. The tube, C D, of the anometric apparatus, is, as we have said, to be placed in com- unication with the air-pump and the balloon, by means of the ree way tube and the leaden pipe *a b*. The gas having been moved, the cock communicating with the air-pump is closed, d the difference of level between the two columns of mercury the tubes, A B and C D, is measured by a cathetometer. This fference is the measure of the elastic force of the gas remaining the balloon. A thermometer, T, indicates the temperature at e time of the experiment.

336. The division, *m n*, in the cistern is necessary, in der to prevent air reaching the barometer, in consequence of e great oscillations in the level of the mercury, during experi- nts. The balloon, thus again exhausted, and having the sticity of its residual gas known, is to be closed, removed, ed, and weighed as before.

337. The difference, $P-p$, between the two weighings, repre- nts the weight of the gas, which at 0° C. = 32° F. fills the loon under a pressure equal to the barometric pressure, H, served at the moment of closing the cock, diminished by the stic force, h, of the gas remaining in the balloon after making exhaustion. The weight of the gas at 0° C., and under the mal pressure of 760 mm. (29·922 inches), is obtained by the mula [1]—

$$(P-p) = \frac{760}{H-h}.$$

338. *Bunsen's method.*—The specific gravities of various es are represented by the weights which equal volumes of se gases possess. As the volume occupied by a given weight any gas is dependent upon the variations of the force of vity accompanying change of geographical latitude, or eleva- above the sea's level, all gases, of which the absolute umes are required, must be reduced to the same latitude and ation above the sea, and to the same barometric pressure temperature.

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[1] Greville Williams's *Handbook of Chemical Manipulations.*

339. According to the most accurate experiments, for which we are indebted to the classical labours of Regnault, one gramme of atmospheric air at the level of the sea, in the 45th degree of latitude, at 0° C., and under a pressure of 760 mm. of mercury, occupies a volume of 773·526 cubic centimetres In the latitude 52° 36′—as, for instance, in Berlin,—a gramme of dry air at 0° C., and under a pressure of 760 mm., occupies exactly 773 cubic centimetres.

340. The accuracy of gasometric determinations is seldom so great that the differences resulting from the variation of gravitation extend beyond the limits of the possible observational errors. Hence, excepting in *normal* determinations, where the greatest accuracy is required, the volume of one gramme of dry air at 0° C., and 760 mm. pressure of mercury, may be represented in our latitudes by 773 cubic centimetres; and the specific gravity of a gas may be defined to be the weight in grammes of gas which, under the same conditions, occupies a space of 773 cubic centimetres.

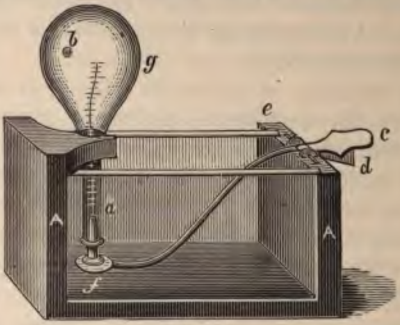

FIG. 56.

341. In cases in which a normal determination is not required, a common light flask, *g*, fig. 56, is employed for measuring the volume of gas of which the specific gravity is to be estimated.

342. The volume of the flask should be about 200 or 300 cubic centimetres, and the neck, *a*, thickened before the gas blow-pipe, must be drawn out so as to have an aperture of the thickness of a straw, into which a glass stopper is ground air-tight by means of emery and turpentine. Through this neck, which is furnished with an etched scale in millimetres, mercury is poured, by means of a funnel reaching to the bottom of the flask, until the whole is filled. As soon as this is accomplished, the flask is transferred, with its mouth downwards, into the mercury trough, AA, and gas is allowed to enter, until the level of mercury in the neck of the flask stands a few millimetres higher than that in the trough. In order to secure the absence of all gaseous impurities, this gas is evolved from as small a vessel as possible, and allowed to enter the flask through a narrow delivery tube, and in the moist state. The gas is dried in the flask itself by a small piece of fused chloride of calcium, *b*, which had previously been made to crystallise on the side of the flask by bringing it in contact with a single drop of water, and alternately heating and cooling the glass. This small piece of chloride of calcium serves also to free the mercury and the sides of the flask from its adhering moisture. In order to be able to close the flask at any time without warming it with the hand, the little lever, *c f*, is employed. On the lower end, *f*, of this lever the stopper is so fastened in a cork that it passes into the neck of the flask without closing it, and the lever is held in its right place by a wedge, *d*, pushed under the finger-plate, *c*. As soon as the apparatus has attained the constant temperature, t, at the barometric pressure, P, the volume, V, of the gas, and the height, p, of the column of mercury rising above the level of the metal in the trough, are observed with the cathetometer-telescope. If the observed volume of gas in cubic centimetres, reduced from a table of capacity, be represented by V_1, this volume, at 0° C., and 760 mm. pressure, becomes, in cubic centimetres,—

$$V_2 = \frac{V_1\,(P-p)}{760\,(1+0\cdot00366\,t)}$$

343. It is now only necessary to determine the weight, G_2, of this volume, V_2. This is obtained in the following manner:— The wedge, *d*, is taken away; the flask is thereby closed, and by withdrawing the pin, *e*, it can then be removed, together with the lever, *c f*, from the trough. After having been most carefully freed from all adhering matters, and having attained the temperature, t_1, of the balance, by the pressure P_1, the flask can be weighed. Let G represent the weight in grammes thus found. The glass stopper is now removed, and replaced by a caoutchouc tube, connected with a drying tube. *b*, fig. 57. The apparatus thus arranged is placed under the receiver of an air-pump, and the air so often withdrawn and admitted until all the gas has been replaced by dry air. If this weight amounts to G_1 grammes, the weight, G_2, of the volume of gas, V_2, measured in the flask, is equal to—

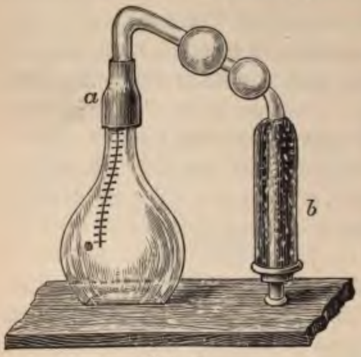

FIG. 57.

$$G_2 = G - G_1 + \frac{V_1\,P_1}{773 \times 760 \times (1 + 0{\cdot}00366\,t_1)}$$ [1]

344. From this value, G_2, the specific gravity is obtained by the help of the following formula :—[2]

$$S = 773\,\frac{G_2}{V_2}$$

Exercises for Calculating by Exact Methods the Weight and Specific Gravity of Gases.

205. Calculate the weight of one litre of dry air at 0° C. and 760 mm. from the following determination by Regnault :—

Globe full of air, and surrounded by ice.

Height of barometer at the time of closing the stopcock	761·19 mm.
Weight added to globe to equipoise it in balance.	1·487 gram.

Globe exhausted of air, and surrounded by ice.

Tension of air remaining in globe, as indicated by the manometer at the moment of closing the stopcock	8·43 mm.
Weight required for equipoise	14·141 gram.

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[1] The unequal specific gravity of the glass and mercury on the one hand, and the metal weights on the other, is not considered in this calculation, as the inaccuracy thus introduced is inconsiderable, in comparison with the observational errors.

[2] From Bunsen's *Gasometry*, translated by Sir H. Roscoe.

206. Calculate the weight of one litre of hydrogen at 0° C. and 760 mm. from the following determinations of Regnault :—

Globe full of gas, and surrounded by ice.

Height of barometer at the time of closing the stopcock	756·16 mm.
Weight added to globe to equipoise it in balance.	13·301 gram.

Globe exhausted of gas, and surrounded by ice.

Tension of gas remaining in globe, as indicated by the manometer at the moment of closing the stopcock	3·40 mm.
Weight required for equipoise	14·1785 gr.

207. Calculate the weight of one litre of carbon dioxide at 0° C. and 760 mm. from the following determinations by Regnault :—

Globe full of gas, and surrounded by ice.

Height of barometer at the time of closing the stopcock	763·04 mm.
Weight added to globe to equipoise it in balance.	0·6335 gr.

Globe exhausted of gas, and surrounded by ice.

Tension of gas remaining in globe, as indicated by the manometer at the moment of closing the stopcock	1·57 mm.
Weight required for equipoise	20·211 gram.

208. The following observations were made in determining the vapour lensity of gaseous bromide of methyl by Bunsen's method. From these lata calculate the specific gravity of the gas.

Volume of gas	42·19 c.c.
Temperature of room	16·8° C.
Height of barometer.	746·4 mm.
Height of the column of mercury above that in the trough	24·3 mm.
Weight of flask and gas	7·9465 gram.
Weight of flask and air	7·8397 ,,
Temperature in balance-case	6·2° C.
Height of barometer at the time of weighing	742·1 mm.

209. The following observations were made in determining the density f a gaseous mixture by Bunsen's method. Calculate the density from hese data.

Volume of gas	211·23 c.c.
Temperature of room	6·2° C.
Height of barometer	760·9 mm.
Height of the column of mercury above that in trough	19·4 mm.
Weight of flask and gas	54·6213 gram.
Weight of flask and air	54·4838 ,,
Temperature in balance-case	7·9° C.
Height of barometer at the time of weighing	760·9 mm.

210. The following observations were made in determining the density of a gas by Bunsen's method. Calculate from these data the density of the gas.

Volume of gas	210·2 c.c.
Temperature of room	19·7° C.
Height of barometer	748·0 mm.
Height of the column of mercury above that in the trough	17·2 mm.
Weight of flask filled with gas	53·591 gram.
Weight of flask filled with air	53·5775 ,,
Temperature in balance-case	20·5° C.
Height of barometer at the time of weighing	·748 mm.

Examination Questions.

211. What is meant by specific weight?

212. Explain why a body weighs more in air than in water.

213. Explain precisely what is meant when it is said that gold is heavier than iron.

214. Explain the use of a table of specific gravities in calculating the weight of assigned volumes of the various kinds of substances; state the principles of the calculation.

215. Explain the method of determining the specific gravity—

1. Of a solid heavier than water; in a state of powder, and also in the mass.
2. Of a solid lighter than water.
3. Of a solid soluble in water.

And state the precautions required to obtain accurate results.

216. Heiro, a king of Syracuse, gave a goldsmith a certain weight of gold to make into a crown. When the king received the crown, he found it weighed exactly the same weight as the gold which he had given to the goldsmith; but he suspected that part of the gold had been replaced by an inferior metal; Archimedes discovered this to be the case, and he ascertained the exact amount of the inferior metal in the crown; state how he could ascertain the amount.

345. *Sp. gr. of vapours.*—We shall limit our description to the means for determining the density or sp. gr. of vapours to the method invented by Victor and Carl Meyer; the two oldest methods Dumas' and Gay Lussac's are now no longer employed; and we refer those of our readers who wish to become acquainted with Hofmann's method to vol. iii. of Roscoe and Schorlemmer's Treatise on Chemistry.

346. V. and C. Meyer's process is suitable for very wide ranges of temperature: it has also the merit of requiring very little substance; it takes but little time, and does not require either the exact temperature of the bath or the volume of the vapour flask to be known. The only observation which is

required is the volume of the vapour in the form of its equal volume of air, measured at the temperature of the air.

347. The apparatus is shown in fig. 58. It consists of the boiling-flask *c*, in which is placed the cylindrical glass vessel *b*, having a capacity of about 100 c.c., with a height of 200 mm. to which is fused a tube 600 mm. long, and 6mm. diameter, and terminating above in a thimble-shaped enlargement, which is closed by a caoutchouc stopper, *d*. A lateral tube, *a*, of not more than 1 mm. internal diameter, and 140 mm. long, is so bent as to carry the air expelled by the vapour into the graduated tube, *e*, which stands in the water contained in the pneumatic trough, *f*. Before commencing the operation, a little recently ignited asbestos is placed at the bottom of *b*, to prevent its fracture when the tube containing the substance to be operated upon is dropped in. In the bulb of the flask *c* is placed the fluid or substance by the vapour of which the vessel *b* is to be heated.[1] The experiment is commenced by heating, by means of the vapour in the flask *c*, the vessel *b*, containing as yet only the asbestos, the tube being closed by the stopper *d*; and the gas delivery tube must dip into the water in the trough, but must not yet be placed beneath the graduated tube. As soon as the temperature becomes constant, and when, therefore, no further evolution of air is observed, the stopper *d* is quickly removed, and a weighed quantity of the substance (such in amount that its vapour does not occupy more than half the volume of the vessel *b*) is introduced into that vessel, and the cork quickly replaced; in the act of replacing it a few bubbles of

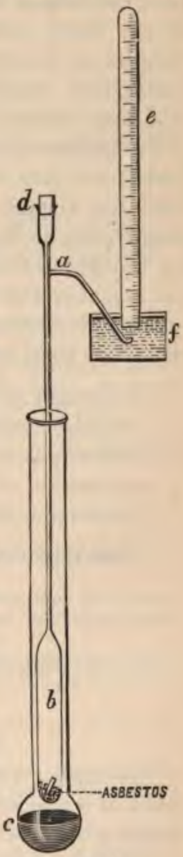

FIG. 58.

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[1] If it be necessary to work at a temperature above 310° C. the heating tube is replaced by a bath of molten lead.

air will escape by the lateral tube *a* ; they must not, however, be allowed to enter the graduated tube *e* ; after they have escaped, the delivery tube must be placed beneath the graduated one, which has been previously filled with water. The substance at once evaporates, and in fifteen seconds displaces its own volume of air, which collects in the tube *e*. As soon as no further bubbles of air are emitted, the operation is finished, and the caoutchouc stopper is instantly removed, to prevent the water in *f* from rushing back into the tube *b*. The lower extremity of the graduated tube is then to be closed with the thumb and transferred into a cylinder of cold water. When a sufficient time has elapsed for the air in the tube to attain the same temperature as the water, the water inside and outside the tube are brought to the same level ; the temperature of the water and the height of the barometer are then accurately taken, and the volume of air in the tube is instantly read off. 'These observations yield sufficient data for the calculation :—

S=weight of substance.
t=temperature of the water.
B=barometric pressure reduced to 0°.
w=tension of vapour of water.
V=volume of air.

'The vapour density is calculated by the formula :

$$\frac{S \times 760\ (1+0{\cdot}003665t)}{(B-w)V \times 0{\cdot}001293}$$

or by collecting the constants :

$$\frac{S(1+0{\cdot}003665t) \times 587780}{(B-W)V}.^{1}$$

Example : 1008 grammes of the vapour of a substance taken at a temperature of 16·5° C., when the barometer stands at 707·5 mm., occupy 22 cc. ; calculate its vapour density.

The barometric pressure of 707·5 mm. at 16·5° C. is equivalent to 705·503 mm. at 0° C. This correction is made from tables constructed for the purpose.[2]

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[1] The student will see that the number 587,780 is obtained by dividing 760 by ·001293.

[2] Owing to the expansion of the mercury and glass tube, which has already been explained, the same pressure will be indicated by different readings on the

We find from the table that the vapour density of water at 16·5° C. is 13·964 mm.

We have therefore the following data :

$$\begin{aligned} S &= \cdot 1008 \\ t &= 16\cdot 5 \\ B &= 705\cdot 503 \\ v &= 22 \\ w &= 13\cdot 964 \end{aligned}$$

substituting these values in the above formula, we obtain for the density required :

$$\frac{\cdot 1008 \times 760 \times (1 + \cdot 003665 \times 16\cdot 5)}{(705\cdot 503 - 13\cdot 964) \times 22 \times \cdot 001293} = \frac{1008 \times 760 \times 1\cdot 06042}{691\cdot 539 \times 22 \times \cdot 001293}$$

$= 4\cdot 13.$

Exercises for finding the Density of a Vapour by Meyer's Process.

217. From the following data determine the vapour density of the organic substance operated on : $S = 0\cdot 0855$ grammes, $t = 16°$ C., $B = 717\cdot 8$ mm., $v = 22$ cc. Barometric correction from Bunsen's tables, 1·971.

218. From the following data determine the vapour density of the organic substance operated on : $S = 0\cdot 0905$ grammes, $t = 17°$ C., $B = 714\cdot 8$ mm., $v = 13\cdot 6$ cc. Barometric correction 2·08.

Examination Questions.

219. Would a lump of cork and a lump of lead, which each weigh 1 lb. in the air, have an equal weight *in vacuo*?

220. If the mercurial barometer stands at 29 inches, and the density of the mercury compared with that of water be 13·57, find what would be the height of a water-barometer.

221. Define exactly what is meant by the specific gravity of a body, and what by its density, and state precisely what is the difference between them.

222. A substance, of which the sp. gr. is 6, weighs 180 grs. in air, and 155 grs. in a liquid ; calculate the sp. gr. of this liquid.

223. A Prussian dollar, made of an alloy of silver and copper, has the sp. gr. 10·05. Determine the amount of silver and copper in it, the sp. gr. of silver being 10·5, that of copper 8·7.

224. What does a barometer measure? At the bottom of a mine a mercurial barometer stands at 77·4 centimetres ; what would be the height of an oil barometer at the same place, the sp. gr. of mercury being 13·596, and that of oil 0·9?

225. Describe the process of determining the sp. gr. of a liquid by

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scale at different temperatures. Hence they are always reduced to the reading at *zero* temperature. To avoid the trouble of calculating, tables have been constructed so that the correction can be immediately ascertained. In the example and exercises the corrections are given.

weighing a solid in it. A piece of glass rod weighs 35 gram. in air, and 21 gram. in water. What is its weight in alcohol of sp. gr. 0·9?

226. Why, in fixing the boiling-point of a mercurial thermometer, must the thermometer not dip into the liquid?

227. Explain why a body weighs less in water than in air. If a piece of metal weigh *in vacuo* 300 grs. more than in water, and 252 grs. more than in spirit, what is the sp. gr. of the spirit?

228. Two similar vessels are filled, one with water and the other with mercury, and are placed over two precisely similar lamps: which would you expect to boil first, and why?

229. Assuming the sp. gr. of mercury to be 13·6, and the mercurial barometer to stand at 29 inches, at what height would an oil barometer stand, the sp. gr. of the oil being 0·845?

230. Why is it easier to raise a bucket filled with water from a well whilst it is in the *water*, than when it is in the air?

231. On the launching of our ironclads, certain facts with regard to them are generally given in the newspapers at the time; for instance, at the time the author was writing this part of the book, the twin-screw armour-plated battle-ship 'Sans Pareil' was launched; it was stated in the newspapers: 'Length, 340 ft.; breadth, 70 ft.; depth, 37 ft. 6 in.; displacement in tons, 10,470.' What is meant by displacement?

232. How would you find the sp. gr. of a piece of sugar?

233. If the sp. gr. of iron be 7·2, what would a cubic foot weigh? And how much less would the cubic foot weigh if it were immersed in water?

CHAPTER VI.

SUBLIMATION — EXPERIMENTS — SIMPLE SOLUTION—EXPERIMENTS—CRYSTALLISATION—EXPERIMENTS—QUESTIONS—CHEMICAL SOLUTION—EXPERIMENTS—PRECIPITATION.

348. IN the conversion of a solid into vapour, if on the condensation of the vapour, it assumes the solid state, the operation is called, as has been already stated, *sublimation*. The apparatus used for subliming substances is generally extremely simple, and is less various than that employed in distillation.

349. We will now give a few practical exercises on the subliming of substances.

Exercises on Sublimation.

234. The apparatus to be employed consists of an evaporating basin, having a beaker inverted over it; the two vessels are joined with paper pasted round; by this arrangement the student will be able to see the subliming as it goes on. The dish ought to be seated in a small sand bath, which is heated by a gas lamp.

235. Place some iodine in the dish, then join a beaker with it as described, and seat the dish in the sand-bath.

236. Perform the same experiment substituting indigo for the iodine.

237. Perform the same experiment substituting mercury biniodide for the iodine.

238. Perform the same experiment substituting naphthaline for the iodine.

350. Many solid bodies, when placed in contact with a liquid, disappear by *dissolving* in the liquid. Example : A lump of sugar placed in water dissolves. The solid is said to be *soluble* in the liquid in which it dissolves, insoluble in that in which it does not dissolve. Ex. : Sugar is soluble in water but insoluble in cold absolute alcohol. The liquid in which the solid dissolves is called the *solvent* or *menstruum.* The more minutely any substance is powdered, the more its solution is facilitated. Solid substances are generally reduced to powder in mortars.

351. Solutions are of two kinds, *simple* and *chemical.* A simple or mechanical solution is the mere dissolving of a solid in a liquid, no chemical change occurring in either ; on the removal, therefore, of the liquid by evaporation, the solid is obtained in its original condition. Common salt dissolved in water affords an illustration of a simple solution. In a chemical solution, the solid and fluid combine together, forming an entirely new substance, from which the original solid and fluid can no longer be extracted by mere mechanical operations. Chalk dissolved in hydrochloric acid affords an example of a chemical solution.

352. The solvent, in a simple solution, cannot dissolve unlimited quantities of the substance to be dissolved ; it can only dissolve certain fixed quantities of the solid, the amount varying with the kind of solid, and the amount of any particular solid varying with the solvent. When the solution contains as great a quantity of the solid matter as it is capable of dissolving, it is said to be *saturated.* A solution is known to be saturated when fresh solid matter of the same sort, on being put into it, remains undissolved.

But as fluids dissolve generally larger quantities of a substance, the higher their temperature, the term *saturated*, as applied to simple solutions, is only relative, and refers invariably to a certain temperature. From the tend [illegible] to diminish the force of cohesion, it naturally results that the solubility of most bodies is increased by heat; thus, 100 parts of water, at 60° F., dissolve 11 of potassium sulphate, and at 212 dissolve 25. At 60° F., 32 parts of dry magnesium sulphate are dissolved by 100 of water, but 74 at 212°. This, however, is not always the case; some bodies, as common salt, are equally soluble in water at all temperatures, whilst in other cases the solubility is greater at particular temperatures than either above or below them. Of this peculiarity, sodium sulphate and nitrate are examples. Thus, 100 parts of water dissolve of dry sodium sulphate, at 32°, 5·02; at 52°, 10·22; at 76°, 28; at 93°, 53; at 122°, 47; and at 212°, 42; the solubility increasing up to 93, and from thence diminishing. 100 parts of water dissolve of sodium nitrate, at 21°, 63; at 32°, 80; at 50°, 23; at 60°, 55; and at 246°, 218 parts. Here the peculiarity is of the opposite kind to what occurs with sodium sulphate; the solubility diminishing up to 50, and from thence progressively increasing. The liquids employed as solvents in simple solutions are water, alcohol, ether, oils, &c. The most important solvent is water; the others are only resorted to when the substance to be dissolved is insoluble in that liquid.

353. Some substances are rendered more soluble in water if the water contains in solution some particular substance; thus lime is much more soluble in solutions of sugar than in pure water.

354. When a substance dissolves in more than one solvent, as iodine in water, alcohol, carbon disulphide, &c., it does not dissolve in the *same* quantities in *equal* volumes of the different solvents. When a substance dissolves only in small quantity in a solvent it is said to be sparingly soluble in that liquid; 1 part of iodine, for example, requires 7,000 parts of water to dissolve it at ordinary temperatures. In some cases one solvent will remove a substance from its solution in another solvent in which it is less readily soluble. For this to take place the two liquids must be incapable of mixing together; see Expt. 246. In some cases a portion of the substance dissolved in a liquid is precipitated from that solvent, on the addition of another liquid in which it is less soluble; for this to occur the two liquids must mix together; Expt. 247 illustrates this fact. In some cases the solution of the same substance in different solvents is not of the same colour; the solution of iodine in water, in alcohol, and in carbon disulphide (Expts. 242, 243, and 244) are striking examples of this.

355. The process of solution is conducted either in evaporating dishes, flasks, or test-tubes. The latter are generally employed when the quantity of the solid operated upon is small.

Exercises on Simple or Mechanical Solutions.

239. Make a saturated solution of potassium nitrate (saltpetre) in cold water, by triturating (rubbing) in a mortar an excess, say 1,000 grains, of this powdered salt, with 2,000 grains of water; after the liquid has become saturated, pour 1,000 grains of the clear solution into a flask; heat the solution in the flask until it boils, but in such a way that the temperature of the solution increases slowly; add from time to time (that is to say, as fast as the salt dissolves) weighed quantities (twenty grains at a time) of the nitrate, until the solution is saturated at the boiling-point.

240. Make a saturated solution of potassium chlorate in cold water, in the way previously described, employing 400 grains of the salt for 2,000 of the water. Take 1,000 grains of the clear saturated solution, and apply heat until it boils, in the way previously described, adding during the time the solution is being heated, weighed quantities of the salt, twenty grains at a time, until the solution is saturated at the boiling-point.

241. Make a saturated solution of common salt (sodium chloride) in cold water; employ 1,000 grains of salt for every 2,000 grains of water. Take 1,000 grains of the clear saturated solution, and heat it until it boils, adding during the time the solution is being heated ten grains of the salt; this ought not to dissolve, as common salt is almost equally soluble at all temperatures.

242. Add a little iodine to about half a pint of cold water and shake them well together.

243. Add iodine to about an ounce of alcohol as long as it continues to dissolve.

244. Add a *very little* iodine to about an ounce of carbon disulphide.[1]

245. Observe the different colours of these three solutions of iodine.

246. Add to the aqueous solution of iodine obtained in Experiment 242, a little carbon disulphide, and shake the two liquids together, and then allow the liquid to stand; the carbon disulphide will fall to the bottom, and it will be readily apparent that it has removed all the iodine from its solution in water.

247. Add some water to the alcoholic solution of iodine obtained in Experiment 244; some of the iodine will be precipitated from the solution in the form of a brown powder.

248. Make a saturated solution of lime in cold water; employ five grains of quicklime for 2,000 grains of water. Take 1,000 grains of the clear saturated solution and boil it, which will cause the liquid to become turbid, as lime is less soluble in hot than cold water.

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[1] By dissolving a large quantity of iodine in carbon disulphide, a solution is obtained which is perfectly opaque to rays of light, though it allows heat rays to pass freely.

356. When a saturated solution of a substance, which is more soluble in a hot than a cold liquid, is made at the boiling-point, the proportion which has been dissolved by reason of the temperature separates from the solution when it becomes cold. Thus, if 151 parts of magnesium sulphate be dissolved in 100 parts of boiling water, and allowed to cool to 60°, a quantity of the salt in a crystalline form will be obtained, weighing 86 parts, for at 60° the 100 of water can only dissolve 65 parts of the salt, and the difference between that and the 151, which had been dissolved by the boiling water, must separate from the solution in the solid state. If the body be like common salt, equally soluble in water at all temperatures, the above process cannot be applied, and a quantity of the liquid must be evaporated, an operation now to be explained.

357. Evaporation is an operation which has frequently to be performed for the purpose of obtaining a more *concentrated* solution of the dissolved body; this is effected by volatilising a portion of the *solvent*. Thus, if we have a weak solution of magnesium sulphate, and we required it in a more concentrated state, we should evaporate so much of the water as would render the solution of the necessary degree of concentration.

358. Evaporation is also had recourse to for the purpose of obtaining a solid back from its solution; this is effected of course by vaporising the solvent. Thus, if we desired to obtain sodium chloride from a solution of this salt in water, we should evaporate the water, and thus obtain the salt in the solid state. If the evaporation be conducted slowly, the solid matter will frequently, on being deposited, assume a crystalline form. The operation is then termed *crystallisation*.

359. When a liquid evaporates at common temperature, it is said to undergo spontaneous evaporation.

360. We can frequently separate two or more solid bodies from one another by taking advantage of the different degrees of solubility which different substances possess. Example: Commercial potassium nitrate (saltpetre) always contains variable quantities of potassium chloride and sodium chloride, and from these substances it must be entirely free when used in the manufacture of gunpowder. Now, 100 parts of water at 70° F. dissolve the same quantity (35 parts) of each of these salts; but 100 parts of water at 212° F. dissolve 240 parts of potassium nitrate, 35 parts of sodium chloride, and 55 of potassium chloride. Suppose we had some saltpetre containing, for every 240 parts of that salt, 30 parts of sodium chloride, and the like quantity of potassium chloride; if we were to dissolve it in the proportion of 300 parts of the salt to 100 of boiling water, and let this solution become cold, about 200 parts of the potassium nitrate would crystallise out, whilst the whole amount of the other two salts would remain in solution. These 200 parts of potassium nitrate would be perfectly pure, with the exception of the small quantity of mother-liquor,[1] with which they would be moistened; to free the crystals from this impurity they may be washed with a small quantity of cold water, or better with a saturated solution of pure potassium nitrate.[2] If the mother-liquor, which would contain nearly equal quantities of each of the three salts, were evaporated at 212°, 84 parts of water out of every 100 could be evaporated, without any of the potassium nitrate being deposited, whereas the expulsion of so much water would cause 25 parts of sodium chloride, and 22 of potassium chloride, to separate. If the clear solution (that is, free from the chlorides which have crystallised out), when it has reached this degree of concentration, were drawn off into another vessel and allowed to cool, a further crop of crystals of potassium nitrate, nearly pure, would be obtained; they might easily be rendered perfectly pure by washing them in the way just stated.

Exercises on Crystallisation.

249. Allow the boiling saturated solutions obtained in Experiments 239, 240, to become cold; as these solutions cool, the salt they contain will begin to crystallise out, and when the solutions have become quite cold all the salt dissolved by warming the solution will in both cases crystallise out.

250. Dissolve some cupric nitrate in water, then evaporate the solution until it reaches the crystallising point, then place it aside to cool; when it has ceased to deposit crystals, pour the liquid from them, and then dry them between blotting-paper, and afterwards place them in a bottle. Evaporate the solution again, and so obtain a further crop of crystals.

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[1] The liquor from which the crystals have separated.

[2] After water has been saturated with one substance, although it cannot dissolve a further quantity of that one, it can still dissolve other substances.

257. Dissolve some magnesia in dilute hydrochloric acid.

258. Dissolve some iron filings in dilute sulphuric acid, one part of strong acid and three of water.

259. Mix equal parts of common salt, calcium carbonate (chalk), and sand together. The first is soluble in water, the second is insoluble in water, but soluble in hydrochloric acid; the last is insoluble both in water and acids. Take a few grains of the mixture, and separate the three substances from each other by dissolving out the salt by means of water,—then the chalk, by means of dilute hydrochloric acid, so that finally no other substance but the sand remains undissolved.[1]

363. When a substance in *solution* is added to another in *solution*, and a substance is formed by the union of one or more of the constituents of the two substances, which is *insoluble* in the liquid in which the two substances were dissolved, the operation is called *precipitation*, the insoluble substance is called the *precipitate*, and the substance which is added to the other substance is called the *reagent* or *precipitant*. The precipitate generally falls to the bottom of the liquid, with *different degrees* of readiness, owing to its being *specifically heavier* than the liquid; but in some cases, owing to its being *specifically lighter* than the liquid, it floats on the surface of the liquid; in each case, however, it is called the precipitate.

364. Precipitates are classified, according to their appearances, into crystalline, pulverulent, flocculent, curdy, and gelatinous. The terms *turbid* and turbidity are applied when the precipitate is so small that it cannot be distinguished, except by impairing the transparency of the fluid.

365. Precipitation is an operation which is constantly practised in the preparation of substances in the chemical manufactory as well as in the laboratory. We also resort to it in the laboratory, for the purpose of detecting and separating substances from one another. Thus, if we had a solution which might contain some compound of baryta and lime,—to ascertain whether baryta was present, and if it was, to separate it from the solution, before ascertaining whether lime was likewise present, we might add a soluble chromate; if baryta was present, the chromic acid would combine with it, and as barium chromate is insoluble in water, it would precipitate, whilst calcium chromate, being soluble, would remain in solution. If we were now to add to the clear filtered solution some s oxalate—calcium oxalate, being insoluble in water, would pi tate. If the chromate pro no precipitate, there be no baryta ; if the o produced no precipitate having separated the bar present, lime must be al

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[1] The student, when he has completed the study of this chapter, must pass on to par. 478, Chap. XI., and complete the study of that chapter.

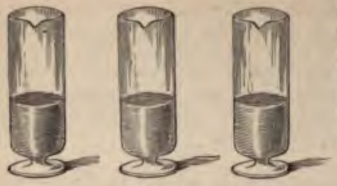

FIG. 59.

366. The separatio precipitates from the liq with a few exceptions, much assisted by the application o and agitation. The operation is conducted in test-tub used as an analytical process ; at the lecture table it mu performed in flasks, if heat has to be applied ; if not, in v of the description illustrated in fig. 59.

367. The student has already been made acquainted filtration ; and that by the aid of filtration we can sepa substance *insoluble* in a from other substances th soluble in the liquid ; bu mere filtration does not pletely separate a preci from the soluble substa inasmuch as it remains r ened with the liquid in they are dissolved ; in chemical operations it h be completely freed from substances ; this is effect washing the precipitate distilled water. It is u washed by means of a bottle ; the usual arrange of the bottle is that sho fig. 60, but there are other of it ; one we have give 30, par. 188) ; the one s here consists, as seen by the figure, of a flask in which are by means of a perforated cork, two bent tubes; the operator

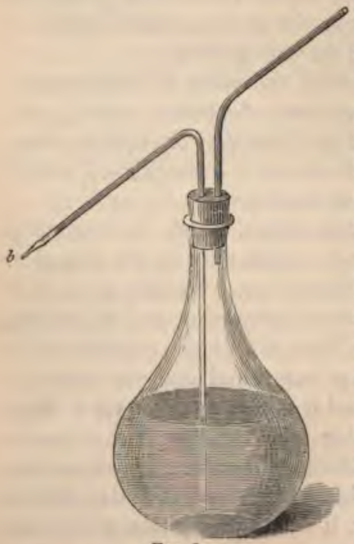

FIG. 60.

through the blow-tube *a*; this forces a fine strong jet of water through the other, *b*, which is directed on the precipitate. If hot water can be used the precipitate is sooner washed free from the extraneous matter than if cold water were used; sometimes other liquids, as alcohol, ether, &c., have to be used in place of water.

CHAPTER VII.

ADHESION—CAPILLARY ACTION—EXPERIMENTS—ADHESIVE ACTION OF CHARCOAL—EXPERIMENTS—DYEING—EXPERIMENTS—QUESTIONS.

368. The cohesive force we have considered; we have now to become acquainted with a series of phenomena, which are considered to be due to another molecular. force, viz. that of *adhesion*.

369. This force may be, and most probably is, merely a modification of the cohesive force; and as the phenomena produced by the operation of these two forces are so similar in some cases, scientific authorities differ as to which of the two forces the *adhering* is due; this difference of opinion arises in the case of the adhering of pieces of the *same* substance, as the adhering of two pieces of lead or two pieces of glass; some authorities consider the action due to the cohesive force, other authorities consider it due to the adhesive force.

370. Whatever difference of opinion there may be as to which of these two forces is to be attributed the adhering of pieces of the *same* substance, there is no difference of opinion as to the adhering of *unlike* substances, as, for example, the writing with chalk, charcoal, plumbago, &c. The adherence of unlike substances is always attributed to the adhesive force.

371. The principal cause of the resistance to motion which is termed *friction* is due to the adhesive attraction exerted between solids. The action of cements is due to adhesion: and that adhesion is exerted between *different kinds* of matter with *varying degrees of force* is seen in the variety of cements required for the variety of materials that have to be united; glue, for example, will join pieces of wood together, but is valueless for joining stone or brickwork together.

372. The action of the adhesive force is not confined to solids, but is exerted, as will be shown, between solids and liquids, solids and gases, between liquids, and the adhesion of gases to liquids.

373. When a piece of glass or wood is dipped into water or oil, the solid when withdrawn comes out wetted, owing to the adhesive attraction between the solid and the liquid; but if the solid is dipped into mercury on its withdrawal no mercury will be adhering to it, owing to the cohesive force in this case exceeding the adhesive one. The *simple* solution of solids in liquids is due to the adhesive force; in these cases the adhesion overcomes the cohesion existing between the particles of the solid; hence, by weakening the cohesion between the particles of a solid, by powdering it or other means, its solution is hastened.

374. From the facts just stated we learn that *a solid body is moistened by a liquid if the cohesion of the liquid is less than the adhesion between it and the solid*; *but if the cohesion is greater than the adhesion*, *the solid will not be moistened by the liquid.*

375. The adhesion of some, and the non-adhesion of other liquids to solids give rise to some important effects due to *capillary* attraction, so called because they are best seen and studied in tubes the bore of which is not greater than the diameter of a hair; the Latin word for hair being *capillus.*

376. The surface of a liquid in equilibrium is a *horizontal plane*, but in contact with solids *capillarity* causes it to be more or less *curved*; being *concave* if the liquid adheres to (wets) the solid, *convex* if it does not. Fig. 61 represents two glass tubes of different diameters inserted in a liquid which wets the sides of the glass. By the aid of the figure we learn two facts: 1st, that the liquid is curved, and that the curve is concave, which is called the *concave meniscus*; 2nd, that the liquid stands higher in the narrower tube.

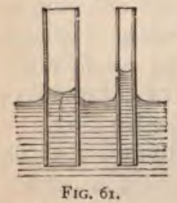
FIG. 61.

Fig. 62 represents the insertion of two glass tubes of different diameters in mercury; the glass is not moistened by the liquid metal ; the mercury, in consequence of the non-adhesion, is depressed at the sides of the glass, and below the level of the exterior mercury, and the curve is convex, which is called the *convex meniscus* ; and the depression, it will be observed, is the greater in the narrower tube.

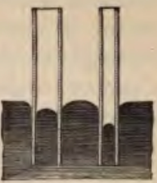
FIG. 62.

377. The following *law* of *diameters* has been arrived at—Capillary elevations and depressions, when all other circumstances are the same, are inversely proportional to the diameters of the tubes.

378. The elevation of the column of liquid in tubes of equal diameter varies with the nature of the liquid, the variation depending partly on the difference of cohesion between the particles of the liquid, partly upon the difference of adhesion between the liquid and the glass. In consequence of the decrease of both these forces by heat, the height of the column diminishes as the temperature rises.

Experiments on Capillary Attraction and Depression.

260. Insert several glass tubes open at both ends, and of different bores, in a vessel of water ; and observe whether the height at which the water stands and its curvature corresponds with what has been stated in the text.

261. Insert tubes of similar sized bores in alcohol and note the results.

262. Insert tubes of similar sized bores in mercury and note the different effects.

379. Both the elevation, and the depression of liquids in tubes above and below that of the liquids surrounding the tubes, appears at first sight to be opposed to the law of gravitation ; but these facts are really no contravention of that force ; as we have learned that, although the molecules of gases repel each other, the atmosphere, nevertheless, cannot indefinitely expand ; it ceases to expand when its elastic force and the attraction of gravitation are balanced ; in like manner, the adhesion between a liquid and a solid causes the liquid to ascend in a tube to a height above the liquid surrounding the tube, and were the adhesive force unrestrained the liquid might ascend in the tube to any height ; but the height is *limited* by the action of gravity, the pressure of the atmosphere, and the

cohesion existing between the particles of the liquid. The influence that the pressure of the air exercises is illustrated by placing in water a capillary glass tube *shorter* in length than the height a column of water would ascend to if the tube were of proper length; the water will ascend to the *top* of the tube, but will not *overflow*, and the curvature of the liquid which was concave becomes convex when it reaches the top.

380. In the case of capillary depression, the liquid sinks as much below the original level as will bring the depressed column of liquid into equilibrium with the capillary repulsion.

381. In noticing the corrections that require to be made for reading accurately the barometric height, we stated that the capillary action between the glass and mercury was one, and that we should notice it in this chapter on adhesion. The capillary depression, we now know, will vary with the diameter of the tube; in tubes of large bore the depression is inappreciable; large-bore tubes are therefore always employed in constructing important standard barometers. In tubes of small bore the depression is sufficiently great to render it necessary to make a correction; the following table will therefore be found useful, as it shows the difference in depression in tubes of different diameters.

Diameter of Tube. Hundredths of an Inch	Depression. Hundredths of an Inch	Diameter of Tube. Hundredths of an Inch	Depression. Hundredths of an Inch
8	17	26	4·0
10	14	28	3·5
12	11	30	3·1
14	9	32	2·8
16	8	34	2·5
18	7	36	2·2
20	6	38	1·9
22	5·2	40	1·7
24	4·5		

382. In reading off the level of mercury in a barometer, or in a eudiometer, the height of the metal should be taken from the convexity of the curve; but in estimating the volume of a liquid which wets the surface of the glass the determination should always be made from the bottom of the curve.

383. If in place of the tubes we employ two glass plates, nd so place them in water, or other adhesive fluid, that one nd of each touches the other, their sides being at a small listance asunder, the highest point of the liquid will be where he plates are nearest together. If the plates are immersed in mercury, a corresponding depression is produced.

384. Everyone is familiar with capillary attraction in a variety of forms:—the ascent of oil, or melted tallow or wax, in the wicks of lamps and candles; the absorption of liquids by sponge, blotting-paper, &c.; the swelling of wood in damp weather; the dampness in the basement storey of honses; and the filtering of liquids through any porous vessel that may contain them, will immediately recur to the student as familiar examples of capillarity. There is one example which at first appears rather perplexing: water will pass through gauze, but mercury will not, unless the gauze is wetted with water, then the mercury passes through; the reason is this, the cohesion existing between the particles of the mercury preponderates over the adhesion of the gauze for the mercury, but this is overcome when the gauze is wetted, and of course the passage of the mercury through the pores is aided by gravity.

385. Immense force can be developed by capillary action: large blocks of stone, for example, can be split by the insertion of wedges of porous wood, which are then saturated with water; the absorption causes the wood to expand, and by the expansion the rock is split. Considerable weights may be raised in the following manner: let one end of a rope be fixed at a point, and the other end to a weight vertically below the point; the rope is then stretched tight, and afterwards wetted; the wetting causes the rope to swell, the swelling shortens its length, and owing to the shortening the weight is raised.

386. The adhesion of a solid body for other solid bodies is sometimes so extremely powerful that it can remove them from their *solutions* in water and other liquids. This is notably the case with charcoal; charcoal is obtained from both vegetable and animal substances; animal charcoal is more powerful than vegetable charcoal, and it is usually prepared from bones by exposing them to a high temperature in closed vessels; by thus excluding them from the air during the time they are heated the gelatinous tissue of the bone becomes carbonised or

Examination Questions.

272. Four cistern-barometers, the diameters of the tubes of which are $\frac{1}{8}$, $\frac{1}{4}$, $\frac{3}{8}$, and $\frac{1}{2}$ inch respectively, are placed side by side. Will their readings be all alike? Give reasons for your answer.

273. A platinum dish containing ether is set in a small pool of water under the receiver of an air-pump, which is then worked; the ether boils, the water freezes, and the ice is full of bubbles: explain these facts.

274. When a bottle of soda-water is opened, the temperature of the liquid falls: explain the reason.

275. Two similar silver vessels containing, one an ounce of water, and the other an ounce of oil, both at 60° C., are taken into a cold room; which will cool the quicker? If one falls to 59° in three minutes, and the other takes 11 minutes to fall to 59°, what is the ratio of the two specific heats?

CHAPTER VIII.

CHEMICAL AFFINITY—ALL SUBSTANCES EITHER ELEMENTARY OR COMPOUND—DISTINCTION BETWEEN ELEMENTARY AND COMPOUND SUBSTANCES—CLASSIFICATION OF THE ELEMENTS—EXERCISES—ELEMENTS CAN COMBINE TOGETHER—EXPERIMENTS—CHARACTERS OF CHEMICAL ATTRACTION—THE DIFFERENCES BETWEEN A MIXTURE OF SUBSTANCES AND A CHEMICAL COMPOUND—CHANGES WHICH ATTEND CHEMICAL COMBINATION—EXPERIMENTS—THE DIFFERENT ELEMENTS DO NOT UNITE WITH EQUAL FACILITY—EXPERIMENTS—TEMPERATURE AT WHICH COMBINATION TAKES PLACE—EXPERIMENTS—SUSPENSION OF CHEMICAL ACTION BY DEPRESSION OF TEMPERATURE—EXPERIMENTS—THE TERMS EMPLOYED IN ORDINARY LANGUAGE TO EXPRESS THE CHEMICAL UNION OF SUBSTANCES—EXPERIMENTS—DEVELOPMENT OF HEAT BY CHEMICAL ACTION—EXPERIMENTS—INFLUENCE OF PRESSURE ON THE HEAT AND LIGHT OF COMBUSTION—EXPERIMENTS—QUESTIONS.

391. WE have now to become acquainted with the other molecular force, viz. *chemical affinity* or *chemical attraction*, and the *changes* produced on matter by the action of this force. We shall, as we proceed, find that it differs essentially from the other two molecular forces. Substances when they are united together, either by the cohesive or adhesive force, undergo, we have seen, not the least change in their properties by the union; whereas when substances are united together by the *chemical force* they lose, as we shall learn, their *individual* properties, the substances produced by the union being endowed with totally new ones. We will commence as before with what the

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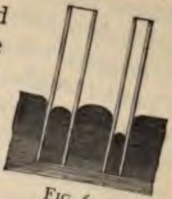

Fig. 62.

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this much admired 'brilliant' is carbon in its purest state; graphite or black-lead, used in the manufacture of pencils, is the other form.[1] It is obtained artificially from several of its compounds, and, according to its source and mode of extraction, it passes under the name of lampblack, bone or ivory black, soot, charcoal, &c.

Table of the more commonly occurring Elements with their Symbols.

Names of the Elements		Symbols	Names of the Elements	Symbols
NON-METALS.			METALS (*continued.*)	
Hydrogen	Gaseous at the common temperature.	H	Calcium	Ca
Oxygen		O	Chromium	Cr
Nitrogen		N	Cobalt	Co
Fluorine		F	Copper (Cuprum)	Cu
Chlorine		Cl	Gold (Aurum)	Au
Bromine fluid		Br	Iron (Ferrum)	Fe
Iodine		I	Lead (Plumbum)	Pb
Carbon	Solid at the common temperature.	C	Lithium[2]	Li
Boron		B	Magnesium	Mg
Silicon		Si	Manganese	Mn
Phosphorus		P	Mercury (Hydrargyrum)	Hg
Sulphur		S	Nickel	Ni
Arsenic		As	Platinum	Pt
METALS			Potassium (Kalium)	K
			Silver (Argentum)	Ag
Aluminium		Al	Sodium (Natrium)	Na
Antimony (Stibium)		Sb	Strontium	Sr
Barium		Ba	Tin (Stannum)	Sn
Bismuth		Bi	Zinc	Zn
Cadmium		Cd		

399. Mercury is the only metal that is liquid at common temperatures; all the other metals are solid bodies. Fluorine is stated in the table to be gaseous at common temperatures, but it is *the* element that is the most difficult to obtain in its *free* or *uncombined* state; it is still doubtful whether it has yet been obtained in its elementary state; and therefore it is still uncertain what its physical state would be if it were obtained uncombined with any other element. We have already noticed

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[1] This name was given to this form of carbon when it was considered, from its appearance, that it was a lead substance.

[2] Lithium is a rare element, but as some of its compounds are extensively used in medicine, we have thought it desirable to introduce it into the table.

that all the gaseous elements have been liquefied; and one of them, oxygen, has been solidified.

400. The most obvious differences between metals and non-metals are physical. The metals in the mass, but not in a finely divided state, reflect light powerfully, which gives them that peculiar appearance termed 'metallic lustre.' Metals are also good conductors of heat and electricity. These three properties are never found associated in the non-metals. A difference in the chemical properties of the compounds the metals and non-metals form with oxygen is noticed in par. 468.

Exercises on Chemical Nomenclature and Notation.

276. Write out the symbols for bromine, boron, barium, and bismuth.

277. What elements do the following symbols denote,—C, Cs, Cd, Cr, Ca, Cl, and Co?

278. What elements do the following symbols denote,—H, Hg, O, N, Pb, Au, Na, Sn, Fe, K, and Ag?

279. Do the following symbols denote compound bodies; if so, what are the elements composing these different compounds—Na, Cl, AgI, KBr, and HCl?

401. Compound substances, we have stated, are formed or produced by the chemical union of two or more elements; the elements must, therefore, be capable of chemically uniting together. That two elements are capable of uniting together and forming a compound substance the following experiments are given as proofs:—

Experiments on the Combination of Two Elements.

280. Place a piece of phosphorus, about half the size of a pea, in the deflagrating spoon (having previously dried it by pressing it gently between folds of blotting paper),[1] (fig. 11); hold the spoon in the flame of a lamp until the phosphorus takes fire, introduce it then into a jar (fig. 12) filled with oxygen.[2] The light produced by the combination of these two elements is so intense that the eye can scarcely bear it. The compound pro-

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[1] The experimenter must dry the phosphorus as expeditiously as possible, and avoid friction, as phosphorus takes fire readily and causes very severe burns; he must also obey the instructions as to the size of the pieces of phosphorus to be used in the experiments.

[2] Oxygen is most conveniently prepared from potassium chlorate ($KClO_3$). For this purpose, powdered potassium chlorate is intimately mixed with one-fifth of its weight of manganese dioxide (MnO_2), and then introduced into a flask, or large test-tube, with a bent tube, as shown in figs. 17 and 18. The flask, or test-tube, may be heated by a gas or spirit lamp—at first moderately,

duced, appears at first under the form of white smoke, which gradually collects into flakes, falls to the bottom of the vessel, and is dissolved in the water. The white compound is called phosphorus pentoxide (P_2O_5).

281. Dry a piece similar in size and in the same way; cut it into small fragments, and introduce them *unignited*, by means of the spoon, into a bottle filled with chlorine.[1] The phosphorus will instantly take fire, burning with a yellowish-green flame. The compound produced is called phosphorus trichloride (PCl_3).

282. Place a few fragments of phosphorus (dried) upon a plate, and throw upon them a small quantity of iodine; the two elements will instantly combine, and form a red solid called phosphorus tri-iodide (PI_3). The combination is attended with evolution of heat and light.

283. Place a piece of roll sulphur about the size of a pea in the cup of the spoon, hold the cup in the flame of a lamp or candle until the sulphur takes fire, and then introduce it into a wide-mouthed bottle or jar filled with oxygen. The sulphur will instantly burst into a brilliant purple flame and rapidly disappear. The two elements form by their union a gaseous compound called sulphur dioxide (SO_2), which is very soluble in water, and has an intensely suffocating odour.

284. Introduce, by means of the spoon, a piece of red-hot charcoal into a jar filled with oxygen; the charcoal enters into a vivid state of combustion. The compound produced by the union of these two non-metals is a gaseous compound called carbon dioxide (CO_2).

285. Coil a thin iron wire round a thin stick so as to bring it into a corkscrew shape, and then draw it off. Tip one end of the wire with sulphur, by immersing it in melted sulphur, the other end being fixed into a cork which fits the mouth of the gas jar. The sulphur having been ignited, the spiral wire must be immediately inserted into a jar filled with oxygen, and the cork is in this case not closely pressed into the mouth of the jar. The sulphur will burst into full flame; the iron, thus becoming heated strongly, is capable of entering into combination with the oxygen; a brilliant light accompanies the union, and the compound produced, an oxide of iron, falls down in brilliant sparks.

286. Put a little powdered antimony into a small fine muslin bag, and shake the bag over a bottle filled with chlorine, in such a manner that the small particles of metal coming through the muslin may fall into the bottle; they will take fire and burn brilliantly, a compound of chlorine

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as the gas, on the first application of heat, comes off very abundantly and rapidly; the heat is gradually increased as the evolution of gas decreases. The gas may be collected over water in the pneumatic trough.

[1] Chlorine is prepared in the following way:—One part of finely-powdered manganese dioxide (MnO_2) is placed in a capacious flask, and four parts of hydrogen chloride (hydrochloric acid, HCl) are added; the flask is fitted with a tube bent at right angles—the shorter limb is fitted into the neck of the bottle by means of a perforated cork, and the longer one is conducted to the bottom of a dry, wide-mouthed, stoppered bottle. The flask is gently warmed, and as soon as the gas begins to be evolved, the heat is removed until it commences to be slowly evolved, when it is again applied. The gas being much heavier than air collects in the bottle by displacing the air; as soon as the bottle is filled, which is easily seen by the colour, it is removed, and another substituted; the full bottle is immediately stoppered, the stopper being first slightly greased.

and antimony being produced. The experiment may be varied by substituting powdered arsenic for the antimony. In both cases the experimenter must be very careful not to inhale any of the vapour, and the apparatus, as soon as the experiment is completed, ought to be removed from the room. If copper or gold leaf be introduced into chlorine, the non-metal and metal combine, and the combination is attended, as in the two preceding cases, with evolution of heat and light.

402. It is not intended that the teacher should explain how the different elements employed in the preceding lesson are obtained from their compounds ; to do so would be to derange entirely the plan of the work. The object of the lesson is not to teach the learner how the different elements are obtained from their compounds, but what is much simpler, viz. that two elements can combine together. The self-instructed student is requested to prepare the different gases without attempting to understand the mode by which they are obtained from their different compounds. The methods for extracting the elements from their compounds form the subject of a future lesson.

403. The student need not trouble himself to remember the names of the compounds formed in the different experiments, or their symbols ; as chemical notation and nomenclature are taught in a series of systematic lessons farther on.

404. The collection and transference of the gases, the vessels in which they are collected, as well as the vessels in which they are generated, are described at the end of Chapter I.

405. It may be stated in general terms, that all chemical experiments are either experiments of *combination* or experiments of *decomposition*, or they are a combination of the two ; the student ought therefore to remember that, however complex any chemical experiment he witnesses may appear, it must either be an experiment of combination, which is termed a *synthetical* experiment, or an experiment of decomposition, which is termed an *analytical* experiment, or it must be both a synthetical and an analytical experiment. No *chemical compound* is composed of all or a majority of the elements ; a large proportion are only made up of some two, three, or four. The elementary substances composing a compound are called its *constituents* or *components*.

406.—The last course of experiments was given to teach the learner the important fact that two elements can combine together, and that a compound substance is produced by their union. One of the experiments must be repeated before we attempt to extract any further information from them ; suppose we repeat the eighth, which was the most brilliant.

287. Burn a piece of phosphorus in a jar of oxygen, in the way described in Experiment 280.

407. Chemical attraction differs from the attractions of cohesion and adhesion in a great many important particulars. When substances combine by means of chemical attraction, they *lose their individual properties* so long as they remain *united* —the compound produced by the union possessing properties perfectly distinct from those of its constituents. If we contrast the properties of the compound we have just formed by the combination of phosphorus and oxygen, with the properties of its constituents, we shall see that this is the case. First of all, the compound we have just formed is a solid body, and therefore oxygen, which, in its uncombined state, is a gas, has become transformed into a solid by its union with phosphorus. Then again, the compound is very soluble in water, and therefore differs from phosphorus, which is entirely insoluble in that liquid, and from its other constituent, oxygen, which is only very slightly soluble ; and it possesses other properties which neither of its constituents possesses.

408. Although, in all the experimental exercises, the student's attention will be drawn to the fact that substances lose their individual properties when they combine together, and that the compound formed assumes properties peculiar to itself, yet it may be as well, whilst we are alluding solely to this particular effect of chemical combination, to adduce a few more examples. Let us take, as our first example, the compound body water. One would not suppose, from its appearance, that it is a compound body—much less that it is composed of two gases, oxygen and hydrogen. One of them, hydrogen, is one of the most inflammable substances in nature, and yet water cannot be set on fire ; oxygen, on the other hand, enables bodies to burn with great brilliancy, whilst water extinguishes combustion. The pungent substance ammonia is com-

posed of the two gases, hydrogen and nitrogen, both destitute of odour. The constituents of sugar are carbon, hydrogen, and oxygen. The first of these bodies is a solid, which has never in its elementary state been dissolved in any liquid; the other two components of sugar are colourless gases,—yet the compound produced by their union is a white, solid substance, which is very soluble in water and has a sweet taste. The compound resulting from the union of the two solid substances carbon and sulphur, is a volatile, colourless liquid, so volatile that it even evaporates in stoppered bottles, unless covered with a layer of water.[1] Hydrogen chloride (muriatic acid) when applied to vegetable substances that are *blue* in colour, changes the colour to *red*; if ammonia be applied to the substance thus *reddened* the *blue* colour is restored; ammonia has, as stated above, a pungent odour; and these two substances are gaseous bodies; when they are brought together they chemically unite, the compound produced by the union being a white solid substance, which does not change or alter colours, and is destitute of odour. As the properties of the two substances are, by the union, destroyed so long as they remain united, they are said to have neutralised each other. The leading circumstance which characterises chemical combination is, as these examples sufficiently illustrate, the loss of properties experienced by the combining substances, the compounds produced being endowed with entirely new ones.

409. But chemical affinity differs from the attractions of cohesion and adhesion in another respect; affinity causes substances to unite in certain *definite proportions by weight*, whereas substances cohere or adhere in any proportions. The full consideration of this peculiarity of chemical affinity forms the subject of a future lesson.

410. From what has already been stated, the learner will be able to understand the difference between a mere *mixture of substances* and a *chemical compound*; in a *mere mixture of substances* the different ingredients composing the mixture do not lose their individual properties, but retain them unaltered and unimpaired, and they can be mixed together in any proportions by weight. But when substances *chemically* combine, they *lose* their *individual* properties, and the compound produced is *endowed* with entirely new properties ; and when substances chemically unite, they can only do so in certain definite proportions by weight. Ex.: 1st. Gunpowder is a mere mechanical mixture of charcoal (carbon), sulphur, and potassium nitrate (saltpetre) ; but when heat is applied to the mixture these substances disappear, for they enter into chemical union, and most of the compounds produced by the substances are gases. 2nd. If sand, lime, and potash are merely mixed together, the three substances each retain their properties unaltered and unimpaired ; but when they are chemically combined we have the substance produced known as glass. 3rd. 'If we mix together iron and sulphur in the finest powder, in the proportions in which we know they will chemically combine to form iron pyrites or ferric disulphide, namely, 47 parts of iron and 53 of sulphur, they will remain a mere mixture, each with its physical properties unaltered. Owing to the presence of iron he powder will have *magnetic* properties, and when placed in water the iron will rust, and remove oxygen from a vessel of air. If a magnet be drawn over the powder the iron will be emoved and the sulphur remain. When the sulphur and iron are, however, chemically united in the proportions above mentioned, as in iron pyrites, they will be found to have lost heir characteristic physical and chemical properties. The ompound has no magnetic properties ; it is not, like sulphur, endered *electric* by slight friction, and in the state of fine owder it does not remove oxygen when placed in a jar of air ver water. A magnet drawn over this powder produces no ffect upon it : the iron is not separated from the sulphur ; we ave, in fact, an entire change of properties, and the new proerties acquired are retained so long as the two elements are hemically combined. By the aid of the chemical force the wo elements may be separated and obtained in their elemen-ary state. The iron and sulphur will then be found to have e-acquired all the properties, physical and chemical, which hey had lost as the result of their combination.'

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[1] Carbon disulphide being heavier, and not miscible with water, this latter liquid floats upon its surface.

411. We have inquired as to the cause of substances chemically uniting ; we have now to inquire whether any out-

and antimony being produced. The experiment substituting powdered arsenic for the antimony. Th menter must be very careful not to inhale any apparatus, as soon as the experiment is completed from the room. If copper or gold leaf be introdu non-metal and metal combine, and the combination two preceding cases, with evolution of heat and lig

402. It is not intended that the teacher the different elements employed in the obtained from their compounds: to do so entirely the plan of the work. The object to teach the learner how the different ele from their compounds, but what is much s elements can combine together. The se is requested to prepare the different gases to understand the mode by which they their different compounds. The methods elements from their compounds form the lesson.

403. The student need not trouble hims names of the compounds formed in the or their symbols: as chemical notation is taught in a series of systematic lessons furth

404. The collection and transference of sels in which they are collected, as well as they are generated, are described in the

405. It may be stated in general terms periments are either experiments of *decomposition*, or they are a combination of ought therefore to remember that however experiment he witnesses may appear periment of combination, which is termed ment, or an experiment of decomposition *analytical* experiment, or it must be both analytical experiment. No *element* all or a majority of the elements made up of some two, three, or fo stances composing a compound are called *ximates.*

406.—The last course of experiments wi

290. Dissolve some ferrous sulphate in water, and add to the solution ome ammonium sulphide; the sulphur and iron will unite and form a olid compound (ferrous sulphide), which is insoluble in water, and of a *lack colour.*

291. Dissolve some zinc sulphate in water, and add to the solution ome ammonium sulphide; the sulphur and zinc will unite and form a solid ompound (zinc sulphide), which is insoluble in water, and of a *white olour.*

413. The preceding experiments are likewise examples of recipitation, which is a *change of form*; the compounds that vere formed, as the learner saw, were insoluble in the liquid in which their constituents were dissolved; they therefore fell r were precipitated to the bottom of the vessel. Chemical ombination is, therefore, sometimes attended with a change f form; the last experiments were examples of this in one irection; the following experiments are examples of it in other orms.[1]

Further experiments showing that chemical combination is sometimes attended with a change of form.

292. Place some fragments of marble or chalk in a deep glass vessel, add some water, and finally some hydrochloric acid; effervescence will be produced on the addition of the acid; effervescence is caused by a gas n its passage through a liquid; the gas (carbon dioxide) in this instance s liberated from the marble by the hydrochloric acid.

When the evolution of gas is instantaneous and for a noment, we call it not effervescence but an explosion; as, for nstance, when we apply heat to gunpowder, and convert it into aseous substances.

293. Take two precipitating vessels of equal size, moisten the sides of ne with a few drops of ammonia, the sides of the other with a few drops f hydrochloric acid; ammonia and hydrochloric acid are gaseous bodies, nd they escape from their aqueous solutions when exposed to the air; the tmosphere of the jars becomes therefore filled with these gaseous bodies. The mouths or openings of the two vessels are then to be brought together, nd the two gaseous bodies will combine, and the compound produced by their union will be a white solid body, which will appear like smoke, owing to its finely divided state. By this experiment we learn that gases may roduce by their union *solids.*

414. It has already been stated, and it will be proved experimentally hereafter, that the elements of water are two

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[1] Refer back to par. 363, and read over that par. and pars. 364–366 on precipitation.'

gases, consequently gases sometimes produce by their union *liquids.*

415. Therefore the changes of form or state that attend chemical combination are exceedingly various ; the combination of gases may give rise to liquids or solids ; solids sometimes become liquids, and liquids solids. And we have seen by the experiments that several familiar chemical phenomena, such as detonation, effervescence, and precipitation are owing to these changes.

416. The next question, I think, any learner would ask about chemical combination would be this,—Can every element unite with every other element? Every element has been made to unite with the greater number of the elements, but never with all ; but the student is not to conclude from this that every element cannot unite with every other element. All we are entitled to say is, that it has not as yet been accomplished. Oxygen, for example, has been united with every element but fluorine, and we conclude from analogy that it would unite with this element if they were brought together under the proper conditions.

417. But although every element may be capable of uniting with every other element, they do not all unite with equal facility, nor do they all unite by direct methods, as the student will learn from the following experiments.

Experiments showing that the elements do not all unite with equal facility, and also that they do not all unite by direct combination.

294. Introduce into a jar filled with nitrogen gas[1] a small piece of ignited phosphorus, by means of the deflagrating spoon ; the phosphorus will cease to burn in the nitrogen ; it will on its introduction into the gas be immediately extinguished, because a union of the two elements cannot be effected under these conditions.

295. Introduce into a jar filled with nitrogen a fragment of ignited sulphur, by means of the deflagrating spoon ; the sulphur, like the phosphorus, will cease to burn in the nitrogen, because the union of the two elements, sulphur and nitrogen, cannot be effected under these conditions.

296. If a piece of burning charcoal in a cage of thin copper wire is introduced. when in a vivid state of combustion, into a jar of chlorine, the

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[1] Nitrogen is most easily prepared by inverting a glass jar in water and burning phosphorus in the confined air contained in the jar. The phosphorus and oxygen combine together, and produce phosphorus pentoxide (P_2O_5), which is dissolved by the water, whilst the nitrogen remains behind almost pure.

l will cease to burn, but the copper will take fire, showing the finity of the chlorine for the copper, and its indifference for car-

Place a piece of phosphorus in a small porcelain crucible or small floating in water contained in a soup plate. Kindle the phosphorus, n fully burning, cover it completely with a jar filled with hydrogen an extinguisher. The hydrogen will be inflamed at the mouth of but the phosphorus and it will cease to burn when the jar rests in If the manipulator is not much accustomed to chemical manipula- had better not perform this experiment, as it requires to be care rformed in order to avoid an explosion.

3. The student must not conclude from these experi- that the elements which were brought together will no they were only given to teach him that all the elements unite with the same facility, and that they cannot all be to unite directly. Nitrogen, for example, has only been o combine *directly* with a few elements.

). The combination of phosphorus with oxygen, and iorus with chlorine, will now be contrasted, and the t must observe the different circumstances which mark mbination of phosphorus with these two elements.

iments showing that the temperature at which combination takes place varies with different substances.

Introduce a piece of *ignited* phosphorus into a jar of oxygen, in directed in Experiment 280.

. Introduce a piece of *unignited* phosphorus into a jar of chlorine, way directed in Experiment 281.

0. These two experiments have been contrasted for the of teaching the student the important fact *that the tem-*

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ydrogen may be prepared by acting on zinc or iron with dilute sulphuric ie part by measure of the acid and eight parts by measure of water. paratus employed may be a Woulfe's bottle (fig. 14); clippings of the iaving been introduced into the bottle, sufficient of the dilute acid is in through the funnel-tube until the metal is covered to the depth of an so with it; at all events, enough acid should be added to produce a it not violent effervescence, and as the end of the funnel-tube is covered uid, the gas evolved can only escape through the other tube. The gas ot be collected until the whole of the air in the different parts of the us has been expelled, as a mixture of this gas and air or oxygen is ously explosive. To guard against this, as much of the newly-formed en ought to be allowed to escape uncollected as would fill the gas ap- two or three times; after this it may be collected with perfect safety. the gas begins to be slowly evolved, the addition of a little more acid the funnel-tube causes it to be evolved again as briskly as at the *perature at which combination takes place varies with the different substances.* In order to effect the combination of the oxygen with the phosphorus, the phosphorus had to be introduced into the oxygen in an ignited state ; whereas the chlorine and phosphorus combined at the ordinary temperature of the air. The student may ask, Would not the phosphorus combine with the oxygen, without the aid of heat ? Not to form the compound produced when the phosphorus burns in oxygen gas, unless it be kept in contact with the air for some time. The reason of this is fully explained in paragraph 462 ; but in order to give the learner a convincing proof that a certain temperature is requisite for the combination of substances, and that the temperature varies with the different substances, the following experiment must be made :—

300. Place in a beaker, or other convenient glass vessel, a piece of phosphorus, and half-fill the vessel with water. Fill a bladder, which is fitted with a stopcock and bent tube (fig. 10), with oxygen, and direct a stream of gas upon the phosphorus. When the two elements come in contact, the phosphorus will not burst into flame ; increase the temperature of the water ten degrees, and then direct the stream of oxygen upon the phosphorus. Continue to increase the temperature of the water ten degrees, after each occasion the oxygen has been directed upon the phosphorus, *until* the phosphorus bursts into flame, when the two elements are brought into contact. When this takes place, the temperature of the water will be about 145° F.

421. We will give a few more examples and experiments to prove that the temperatures at which bodies enter into rapid combustion is very various. Ex. : Chlorine, we have seen, combines with phosphorus (Expts. 281 and 299), with antimony, and arsenic (Expt. 286), at the ordinary temperature of the atmosphere, the combination being attended with evolution of light ; but at the temperature of $-106°$ it does not enter into combination with these substances. Phosphuretted hydrogen bursts into flame when it comes into contact with the air at the ordinary temperature. Carbon disulphide ignites when touched with a warm glass rod. If a jet of coal gas be made to impinge on a bar of iron only *visibly red*, the gas will not ignite : it will only be ignited by the metal when it is bright red. Sulphur enters into combination with the oxygen of the air at about 500°, hydrogen enters into combination with it below 600° ; carbon enters into combination with it at about 1000°, and zinc enters into combination at about 1900°. But it must be observed that the temperature at which combination takes place depends greatly upon the state of division of the substance if it be a solid body ; thus iron, lead, and many other substances which will only combine with the oxygen of the air with vivid combustion at very high temperatures when in *masses*, will when in a state of fine powder combine with the oxygen of the air with vivid combustion at the ordinary temperature of the atmosphere. Paradoxical as it may appear, iron and lead when in a state of fine division ignite at lower temperatures than gunpowder, as the following experiment illustrates.

Experiment showing that Iron ignites at a lower temperature than Gunpowder.

301. Place some spirit of wine upon a dinner plate, ignite it, then allow a little gunpowder to fall through the flame. Perform the same experiment, but substitute for the gunpowder iron *in powder*.

422. A certain temperature, then, is requisite for the combination of substances ; and after combination has commenced, *the heat necessary for its continuance is produced by the combination* ; for we have already seen that heat is evolved when two substances combine together.

423. As a certain temperature is necessary for the combination of substances, it follows *that if the heat evolved by substances in the act of combining be conducted away, the combination must cease for want of the necessary amount of heat.* The following experiments are given to teach the student this fact. He must remember what has been stated, that metals are good conductors of heat.

Experiments showing that combination ceases when the heat evolved is conducted away.

302. Form a small metallic spiral (fig. 63), by twisting a piece of copper wire around a pencil ; if it be placed cold over the flame of a wax candle, the flame is extinguished, because the heat evolved by the combining substances is conducted away by the metal, and therefore the combination ceases for want of the necessary amount of heat. The experiment must be repeated, first heating the spiral to redness in the flame of a spirit lamp, and then placing it over the flame. Under the circumstances, the flame is not extinguished.

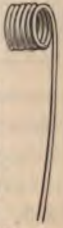

FIG. 63.

424. Inflammable vapour cannot, for the same reason, pass through wire gauze in an ignited state, because the metal conducts away a large portion of heat, and thus reduces the temperature below the point at which combination can take place between the elements of the vapour and the oxygen of the air; the flame ceases therefore as the vapour passes through the orifices.

303. Lay a piece of camphor on some iron wire gauze, and kindle it. The camphor will burn on the surface with a smoky flame, whilst a portion of the melted mass will pass through the gauze. The melted camphor, in its passage from the upper to the under surface of the gauze, will be deprived by the metal of so much heat that, when it comes in contact with the air at the under surface, they will be unable, from the low temperature of the vapour, to enter into a state of combination.

304. Place a piece of wire gauze over a jet from which coal gas is issuing; apply a light to the upper surface of the wire, when the gas will instantly take fire. The metallic gauze may be raised an inch or two above the jet from which the gas is issuing; the gas above the gauze will continue to burn, whilst that below will remain unignited.

425. Sir H. Davy beautifully applied this impenetrability of metallic gauze to ignited vapour, in the construction of a lamp for coal mines, which invention may be truly regarded as 'one of the happiest efforts of his genius.' A gas composed of carbon and hydrogen, called by chemists methane, and by miners fire-damp, often issues in considerable quantities from the fissures in the coal-beds, and gradually mingles with the atmosphere of the mine. This gas not only burns in air, but they likewise form, when mingled together in certain proportions, a most explosive mixture, which on ignition detonates with terrific violence. Davy attempted the task of finding some plan whereby the miner could carry a light into the explosive atmosphere without causing it to ignite, and thus prevent those terrible explosions which were the dread and destruction of the miner. Davy accomplished this apparently impossible task by simply screwing on to the top of a common oil-lamp a cylinder of wire gauze (fig. 64), of about 1,500 orifices in the square inch. The lamp, with this addition, can be carried safely into the most explosive atmospheres; for although the cylinder becomes filled with flame, from the ignition of the explosive atmosphere within the cage, yet the explosive atmo-

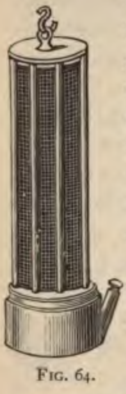
FIG. 64.

flame, which is scarcely visible in the daytime, but is intensely hot. The burning of the hydrogen is caused by its combining with the oxygen of the atmosphere, forming with it water: this is readily proved, for if a dry glass vessel be held over the burning gas, it becomes covered with drops of water. If a glass tube, about three feet long, an inch and a half wide, and open at both ends, be held over the flame, strange musical sounds will be produced, the sharpness varying with the depression or elevation of the tube. This phenomenon is produced by a series of small explosions, which succeed each other so rapidly that they produce a continuous sound. The explosive nature of a mixture of hydrogen and oxygen is better exhibited by filling a *soda-water bottle* two-thirds full of hydrogen and one-third full of oxygen, and then inflaming the gases by applying a light to the mouth of the bottle; the explosion is less violent if atmospheric air be used instead of pure oxygen.[1] If air be employed, five-sixths of it are required for one-sixth of hydrogen; but whether atmospheric air or oxygen be used, it is requisite, before exploding the mixture, to wrap the bottle in a towel.

319. Hydrogen is the lightest body in nature, being fourteen and a half times lighter than air, and sixteen times lighter than oxygen. This property is pleasingly demonstrated by filling a small balloon of gold-beaters' skin with the gas, and allowing it to ascend in the lecture-room. The hydrogen ought not to pass direct from the generating vessel into the balloon; before entering the balloon it ought to be made to traverse a tube filled loosely with fragments of quicklime, which will absorb any sulphuric acid that may have been carried up by the gas, as well as arrest the moisture which accompanies the gas.

320. Fill a soda-water bottle with equal volumes of chlorine and hydrogen, and inflame the mixture by applying a light to the mouth of the bottle,—the explosion is more feeble than with oxygen and hydrogen, and the light is faint. The compound produced is hydrochloric acid. The acid nature of the compound may be demonstrated by applying a piece of blue litmus paper to the bottle.

441. The hydrogen flame, the student will observe, develops little light, although there is a great evolution of heat; indeed, the flame of a mixture of oxygen and hydrogen produces the most intense heat that can be obtained by chemical action.

442. That a solid is necessary for the evolution of light, is beautifully exemplified in the following way:—The combustion of a mixture of oxygen and hydrogen produces the most *intense heat* that can be obtained by chemical action, yet the *light* produced by the combination of the two gases is so small s scarcely to be visible. If we place in this *almost* invisible lame a substance like lime, which will not melt, or undergo any chemical change, a light so brightly luminous will be produced as to rival in intensity the noon-day sun. This light is caused by the solid becoming luminous from being intensely heated.

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[1] The student will see that the difference between combustion and explosion onsists simply in the difference in the rapidity of the combination; in combuson the two substances combine together by degrees, but in explosions they are oth intimately intermingled, and therefore can be made to combine instantaeously.

321. Exhibit the oxyhydrogen limelight,[1] so as to prove to the student that a solid is necessary for the evolution of light.

443. The luminosity of the ordinary hydrogen flame is also greatly increased if a platinum wire coated with a little film of lime, which is obtained by burning upon the wire paper which has been immersed in a solution of calcium chloride, be introduced into the flame.

444. Paper soaked in solution of calcium chloride and burnt in the flame of a spirit lamp, leaves a white network of ashes, which, when held in the feeblest alcohol flame, emits a brilliant light.

445. This method of introducing the solid matter which is to evolve the light, into the burning matter from which the heat proceeds, is one which is never employed for ordinary illumination. The principle upon which all practical methods of illumination proceeding from chemical action are founded is, that the *heat-producing material* and the *incandescent solid* both exist in the substances employed for illumination, and that the burning of the former sets free the latter, yielding at the same time by combustion the necessary degree of heat to render it luminous.

446. All substances employed as illuminating materials are compounds of carbon and hydrogen, and the flame produced by their combustion contains solid particles of carbon, which can be proved by a very simple experiment.

322. Introduce a cold body, such as a plate of metal, or a piece of glass, or even a piece of card, into a luminous flame; it speedily becomes

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[1] The burning of a mixture of oxygen and hydrogen requires great care. The gases ought to be contained in separate holders, and mixed only at the jet; or if they are contained in the same vessel, the tube which conveys them to the jet must be filled with small iron rods, so that in case the flame were to recede, it would be cooled by the rods below the point at which combination takes place between the two gases, so that by no possibility could the flame reach the gaseous mixture in the gas-holder; for a mixture of the two gases, when ignited, explodes with terrific violence.

bination is sufficient to cause the iron and oxygen to combine. A similar plan is adopted for kindling coal fires; for coals, like all other substances, require to be heated to a certain degree before they will burn. This is accomplished, as all know, by the use of the more readily combustible paper and wood.

Experiments showing how the heat necessary for combination to commence is sometimes generated.

306. Coil a thin iron wire round a stick so as to bring it into a corkscrew shape, and then draw it off. Tip one end of the wire with sulphur, by immersing it in melted brimstone, the other end being fixed into a cork which fits the mouth of the gas jar. The sulphur having been ignited, the spiral wire must be immediately inserted into a jar filled with oxygen gas, and the cork is pressed into the mouth of the jar. The sulphur will burst into full flame; the iron, thus becoming heated strongly, is capable of entering into combination with the oxygen; a brilliant light accompanies the union, and the compound produced, an oxide of iron, falls down in brilliant sparks. This experiment is an illustration of the statement made in a preceding page, that the combination of iron with oxygen is sometimes attended with an evolution of both light and heat.

307. We will now make the same experiment, with the exception that instead of plunging the wire into oxygen gas, after the sulphur is ignited, we will allow it to remain in the air, to see if the iron wire will burn in the air: you see that after all the sulphur has entered into combination with the oxygen of the air the combustion ceases, no union taking place between the iron and the oxygen of the air.

429. Atmospheric air is merely a mixture of the gas oxygen, in which the iron burned in Expt. 306, and nitrogen; then why will not the iron burn as well in air as in oxygen gas? Because the nitrogen, which does not enter into combination with the iron (307), or contribute in any way to the combination, absorbs some of the heat evolved by the combination of the sulphur and oxygen, and thus reduces the temperature below the point at which combination will take place between the iron and oxygen. This gaseous body nitrogen acts with respect to the burning of iron in the air in the same way that the metallic spiral acted in Expt. 302, and the wire gauze in Expt. 303. An example of the same kind the student will frequently have seen; he will have seen, for instance, that when glowing coals were placed on considerable masses of metal the combustion was extinguished, simply because the metal conducted away the necessary degree of heat which was required for their combustion.

430. By the non-combustion of iron in atmospheric air and of minute solid particles, which become intensely heated and give out a bright light. The interior part surrounding the wick consists of black, inflammable vapour, which, from having no access to the external air, is incapable of undergoing combustion. The presence of this inflammable vapour may be easily shown by holding obliquely in the flame a glass tube, open at both ends, the lower end of the tube being placed in the centre of the flame, and the upper end reaching beyond its edge. A large quantity of the inflammable vapour will pass up the tube, and may be ignited at the upper end. The cup shape of the candle at the base of the flame, and the tapering form which the flame assumes, are both due to the ascending current of air caused by the combustion of the candle. The flame melts the inner part of the wax or tallow, but the exterior does not melt, on account of the cooling influence of the current: hence the formation of the beautiful cup in candles.

448. In describing the different parts of the flame of a candle, we have drawn attention to the fact, that if the illuminating material be burned in such a manner that its carbon and hydrogen are consumed at the same moment, the largest possible amount of heat is produced, but the amount of light evolved is not greater than that obtained from the burning of pure hydrogen: this is due to the immediate combustion of the carbon, and therefore the absence of any solid matter in the flame, which is necessary, as we have previously stated, for the evolution of light. That illuminating materials derive their light from the separation of their carbon, and that if they be so burnt that this separation does not take place, owing to the carbon being burnt simultaneously with the hydrogen with which it was united, little or no light will be produced, is very simply and very conclusively demonstrated in the following way. Place over a gas jet a chimney, or hollow cylinder, covered over at the top with wire gauze, as shown in fig. 65; when the gas—*i.e.* coal-gas—is turned on, it ascends within the chimney, and mixes with atmospheric air before passing through the gauze. On lighting it above the wire gauze, if the air and gas have been so proportioned that its carbon and hydrogen are consumed at the same time, it will burn with a pale blue flame of no more illuminating power than that of pure hydrogen.[1]

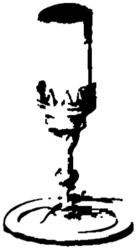

FIG. 65.

449. When the burning body does not exist as a gas naturally, or does not pass into the state of gas at the temperature at which it combines with the other body—oxygen in the air, for

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[1] Bunsen's burner (fig. 19) can be employed in place of this lamp for illustrating the principle stated.

instance—the fire which accompanies the combustion appears as glow, for it is necessary for the production of flame that the burning body should exist naturally in the state of gas, or be converted into gas at or below the temperature at which combination takes place, otherwise no flame results. Well-burned charcoal, and also iron, burns with a steady glow, unattended with flame, as these two elements do not volatilise at the temperature at which their combination takes place. Sulphur, phosphorus, zinc, &c., burn with flame, because they either exist naturally in the form of gas, or they pass into the gaseous state at the temperature at which they enter into combination. Flame is, in fact, produced whenever, by a sufficiently elevated temperature, a continuous supply of vapour or gas is made to combine with oxygen or some other gaseous substance with which it comes in contact. Flame, then, is produced when two vaporous or gaseous substances enter into combination, and the combination is attended with evolution of light and heat; and as combination can only take place at the point of contact, it is the boundary line of the two gases; hence the flame is hollow; that it is hollow may be easily shown by holding a wooden match for a few seconds across the middle of the flame of a spirit-lamp with a large wick; the match will become charred at the edge of the flame, but the intermediate portion will remain uninjured.

450. We have learned, for light to be produced a solid body must be present; but a few remarks have yet to be made on the production of heat by chemical combination. We learned in Chapter II., that, according to the dynamical theory of heat, motion and heat are convertible one into the other; and that the molecules of matter in each of the three states in which matter can exist are, it is supposed, in a constant state of motion; therefore, the molecules of two different bodies must, we are led to suppose, rush together for their atoms to enter into chemical union, and this energy becomes converted by their union into heat.

451. A considerable diminution of temperature occurs in the vessel from which air escapes into a vacuum. Hence, as the gaseous products of combustion increase in volume on the atmosphere becoming rarefied, the temperature of the flame the atmospheric pressure decreases, thus causing the flame to present a gradually increasing surface of contact with the exterior air.

455. We have already stated that the term combustion is employed to express chemical combination when attended with an evolution of light as well as heat. It is not unusual in chemical books to call one of the combining substances the *combustible* body, and the other combining substance the *incombustible* body, or the *supporter of combustion*; thus, in the combination of phosphorus with oxygen the phosphorus would be called the combustible body, and the oxygen the supporter of combustion. We have to prove that these terms are inappropriate, and for this purpose we will commence with the following experiment.

Experiments proving the incorrectness of calling some substances combustible, and others supporters of combustion.

324. Place a piece of sulphur in a long test-tube or flask of hard glass; heat the vessel by means of a lamp until the sulphur melts, and the vessel has become filled with its vapour. Introduce a narrow strip of tinfoil into the sulphur atmosphere; the metal will instantly inflame, a compound of tin and sulphur being formed. The experiment may be varied by employing, instead of the tinfoil, a narrow strip of sheet lead, the combination being attended with evolution of heat and light. If very thin iron and copper wires be twisted into a coil and introduced into an atmosphere of sulphur, they will likewise combine with that metalloid, the combination being attended with evolution of heat and light.

456. The burning of sulphur in oxygen gas, or of tinfoil in sulphur vapour, proves the incorrectness of calling some substances *combustible*, and others *supporters of combustion*; for in the first of these experiments we should have to call the sulphur the combustible substance, whilst in the latter it would be the supporter of combustion. The fact is, that combustion, like all other examples of chemical combination, is really a reciprocal phenomenon, each body, in popular phraseology, burning, or, in chemical language, combining with the other body; and the heat and light are derived from the combination. Oxygen is as capable, for example, of being burnt in an atmosphere of hydrogen or coal-gas as these substances are in one of, or containing, oxygen. If our atmosphere, for instance, had consisted of coal-gas, we might then have burnt oxygen, just as we now burn coal-gas, for the purposes of light and heat. The experiment of burning the oxygen of atmospheric air in an atmosphere of coal-gas may be performed in the following manner :—

325. Fill a bell-jar, that is fitted with a stopcock, with atmospheric air, press it down in water, and open the jet. Bring over the jet a small bottle of coal-gas. Just when the mouth of the bottle is in contact with the jet, apply a light. Immediately depress the bottle so as to bring the jet into the middle of the bottle. Blow out the flame at the mouth of the bottle. The atmospheric air will continue to burn at the jet amid the atmosphere of coal-gas.

457. But the learner will find it stated in many chemical works that this or that substance is not a supporter of combustion. He will find it stated in such works that hydrogen is a combustible substance, but not a supporter of combustion; and in proof of the last assertion, the fact that a lighted taper is extinguished when immersed in an atmosphere of hydrogen or coal-gas is adduced. Now, it would be quite as correct to say that oxygen is not a supporter of combustion, because carbon dioxide cannot be made to burn in an atmosphere of that gas. If one element, B, has already combined with another element, A, in the largest quantity it is capable of uniting with it, the resulting compound cannot, of course, unite with a still further quantity of the element, A. Hence paraffin cannot burn in hydrogen, or carbon dioxide in oxygen; the first, paraffin, being a compound of carbon and hydrogen, cannot unite with a *further quantity* of the latter element; for the same reason, carbon dioxide, being a compound of carbon and oxygen, cannot unite with a *further quantity* of the latter element. To say that an element is not a supporter of combustion is to say that it cannot enter into chemical combination with any other substance, which is incorrect, as every element can combine, *either directly or indirectly*, with some of the other elements, if not with all.

458. The difference in colour different substances impart to a flame will be described in the second part.

459. We will now briefly summarise some of the leading differences between the action of the chemical force and that of cohesion. Like cohesion it acts only at *immeasurably* small distances; but it differs from that force by causing *dissimilar* substances to unite, and they can unite only in certain *definite* and *invariable* proportions by weight: *physical changes* also attend its action, *chemical compounds* having totally different

physical properties to those of their constituents. *Chemical combination* is also attended with an *evolution* of heat, sometimes amounting, we have seen, to vivid combustion; and *chemical decomposition* is attended with an *absorption* of heat. The formation and resolution of chemical compounds are also attended with changes in the *electrical state* of their elements.

The following exercises must now be answered by the student either *vivâ voce* or in writing.

Examination Questions.

326. What is the difference between a mere mixture of substances and a chemical compound?

327. What is the burning of phosphorus, of carbon, of sulphur, of hydrogen in the air due to?

328. State some of the changes which characterise the chemical union of bodies.

329. Name some of the elements which combine at the ordinary temperature of the atmosphere, and name some which require a higher temperature for their combination.

330. Name some of the terms employed in ordinary language to express the combination of substances.

331. State some of the conditions necessary for the combination of phosphorus and oxygen, of phosphorus and chlorine, and the phenomena which accompany the combinations, and the properties of the compounds produced.

332. State the effects which intense heat has upon solid and gaseous bodies, and give examples.

333. Explain the points of difference and agreement between the forces of cohesion and affinity.

334. Name some of the conditions necessary for the combination of antimony and chlorine, and of sulphur and oxygen, and the phenomena which accompany the combinations, and the properties of the compounds produced.

335. Why do substances burn more vividly in pure oxygen than they do in atmospheric air?

336. What is steam? Is it an element or a compound; and can it unite with oxygen?

337. Upon what facts depends the possibility of artificial illumination?

338. What is necessary for the production of flame?

339. Do all the elements unite with equal facility? Support your opinion by examples.

340. Of what elements are all materials employed for illumination composed, and what changes do they undergo when burned in the air?

341. Explain the flame of a candle.

342. Explain very fully the principles on which Davy's Safety Lamp is constructed.

343. Do the light and heat emitted by burning bodies bear any proportion to each other? Confirm the opinion expressed by examples.

344. Supposing a space to be filled with pure coal-gas, would the gas explode on the introduction of a lighted candle? What are the conditions

cessary to an explosion of the gas, and what are the chemical products of e explosion?

345. Two substances are brought together; by what phenomena would u know whether chemical action takes place between them?

346. You now know what compound is formed when phosphorus is rned in air or oxygen; and also what compound is formed when hydron is burned in either of these gaseous substances. What compound or mpounds would you infer would be formed when phosphuretted hydron (PH_3) was burned in air or oxygen?

347. What are the conditions necessary for light to be produced in emical combination?

348. Explain why coal-gas burns quietly when it issues from a gasrner; whilst a mixture of air and coal-gas explodes on being lighted?

349. To what are effervescence and explosion due; and state the difence between the two?

350. When a compound of carbon and hydrogen is burnt in a deficient pply of air, what is the result?

351. A cold plate held above the flame of a candle is bedewed with oisture, and if held in the flame it is covered with soot: explain these cts.

352. State what you know about the point of ignition.

353. Describe Meyer's new method for the determination of vapour nsities.

354. By what characteristics are chemical phenomena distinguished om physical phenomena?

CHAPTER IX.

OMBINATION OF A COMPOUND WITH AN ADDITIONAL QUANTITY OF ONE OF ITS ELEMENTS—EXPERIMENTS—COMBINATION OF A COMPOUND WITH ANOTHER ELEMENT—BASIC OXIDES AND ACID OXIDES—HYDROXIDES—THE ACTION OF BASES AND ACIDS ON TEST-PAPERS—COMBINATION OF COMPOUND SUBSTANCES—SALTS—EXPERIMENTS—THE SOLID STATE LESS FAVOURABLE TO CHEMICAL ACTION THAN THE FLUID STATE—EXPERIMENTS—RELATIVE INTENSITY OF AFFINITY—EXPERIMENT—QUESTIONS—THE PHYSICAL AND CHEMICAL PROPERTIES OF THE MORE COMMONLY OCCURRING ELEMENTS—THEIR EXTRACTION FROM THEIR COMPOUNDS—QUESTIONS.

460. WE have now to inform and demonstrate to the arner that more than one compound can be formed, in many ases, by the union of the same elements; this will at once be ndered easily intelligible to him if we make use of our former omparison; there are different words composed of the same tters; there are, for instance, two different words both comosed of the letters T and O, viz. to and too; the same with

the letters B and E, viz. be and bee; and other examples will at once recur to the student. In like manner, for example, there are two compounds of carbon and oxygen, viz. carbon monoxide, the symbol for which is CO, and carbon dioxide, which, for the present, we will represent by the symbol COO; this latter compound, it will be seen from the symbol, contains double the quantity of oxygen the former contains.

Experiments on the combination of a compound with additional quantity of one of its elements.

355. Place in a test-tube some crystals of oxalic acid, add to them some strong sulphuric acid, and then warm the tube in the flame of a lamp. As soon as effervescence commences, apply a piece of lighted paper, or a splinter of wood, to the mouth of the tube, until the evolved gas takes fire: it burns with a blue flame. The gas is the compound carbon monoxide (CO) we have already alluded to; the heat and light are occasioned by its combining with some of the oxygen of the air, and forming with it carbon dioxide (COO).

461. When a compound contains more proportions than one of one or more of its elements, instead of repeating the symbol in order to indicate the number of proportions, chemists attach a figure to the symbol of the element, or elements, to indicate the number of proportions. The use of the figures in connection with the symbols can be rendered more intelligible perhaps, if we employ our former illustration; the words 'to' and 'too' differ in their orthography, inasmuch as the latter word has one of the letters repeated. From habit we look upon this way of repeating a letter as the only rational one, yet it might be rendered equally intelligible if, instead of repeating the letter, we were to place the figure 2 after it in the following manner, to$_2$. This method chemists employ: thus, for instance, instead of representing carbon dioxide by the symbol COO, they represent it thus—CO_2. We shall hereafter adopt this method for indicating the different proportions of an element or elements in a compound when they are present in more proportions than one.

356. Put into a Woulfe's bottle or flask a few small pieces of metallic copper (copper turnings), and add to the metal, by means of the funnel-tube, equal parts of nitric acid and water; effervescence, without the application of heat, will instantly ensue. On the first evolution of the gas the bottle will be filled with reddish fumes; when they have nearly passed away, the gas may be collected by dipping the exit tube under the mouth

a gas jar or wide-mouthed bottle, which is filled with water, and stand- g inverted on the shelf of the pneumatic trough. This gas is a com- ound of nitrogen and oxygen, and is called nitrogen dioxide (NO).

357. Pass rapidly into a jar half full of this nitrogen dioxide (the vessel ontaining it being inverted on the shelf of the pneumatic trough) a quantity f oxygen; as soon as the two gases are brought together, the jar will be istantly filled with an orange-red gas, which is very soluble in water; this quid will therefore rapidly ascend in the vessel by dissolving the new- ormed gas. The oxygen and nitrogen dioxide unite together, and the ompound produced is this orange-coloured substance named nitrogen etroxide (NO_2). This experiment furnishes a striking example of the dif- erence between a mere mixture of substances and a chemical combination. f the oxygen and nitrogen dioxide had merely mixed, and not chemically nited, these two colourless gases, in whatever proportions they were mixed, onld never have become coloured; neither would the mere mixture of two ases insoluble in water render both of them soluble in that liquid.

358. The experiment may be made in the following manner:—Fill a rge jar with nitrogen dioxide, remove it from the pneumatic trough, and xpose it with its mouth upwards to the air; the compound will instantly ombine with some of the oxygen of the air, and become converted into itrogen tetroxide; immediately on the formation of this latter compound lace the jar, with its mouth downwards, in the water in the trough; the vater will ascend as in the previous experiment.

462. These experiments prove that some compounds have he power of uniting with another quantity of one of their constituents; thus carbon monoxide can combine with another quantity of oxygen, which converts it into another compound, viz. carbon dioxide: in like manner, carbon dioxide can unite with another quantity of carbon and by that union becomes converted into carbon monoxide. This interesting chemical action we frequently witness. The blue flame which is frequently seen on the surface of coal fires, and which is seen issuing from lime-kilns and iron furnaces, is occasioned by the carbon monoxide combining with some of the oxygen of the air, carbon dioxide being formed, as in Experiment 355. The combination is produced in our fires in the following way; the air being in excess at the bottom of the grate, the oxygen unites with the carbon in its highest quantity, forming carbon dioxide; as this gas ascends up through the red-hot fuel it gives up one-half its oxygen to another portion of carbon; the carbon dioxide is thus reduced to carbon monoxide; and carbon monoxide is also formed by the union of another portion of carbon with the oxygen given up by the dioxide. When the carbon monoxide reaches the surface of the fire it combines with some of the oxygen of the air, and becomes converted

into the dioxide, and the blue flame is the effect of the combination.

463. Nitrogen dioxide cannot combine with oxygen if this element be already in a state of chemical combination; it can only appropriate it when it is uncombined; in other words, when it is in a pure state or is one of a mere mixture of gases. As it can abstract oxygen from the atmosphere, we conclude, from this and other facts, that the atmosphere is not a *chemical compound*, but a mere mixture of oxygen and nitrogen.[1]

464. But some compounds can either combine with another proportion of one of their constituents or with some other element. Example: Carbon monoxide can take up an additional quantity of oxygen, which converts it into carbon dioxide; likewise if carbon monoxide and chlorine be mixed together, they combine, a compound called carbon oxychloride being formed. Sulphur dioxide (SO_2) (the compound produced by burning sulphur in oxygen) can combine, under certain circumstances, with another proportion of oxygen, which converts it into another compound, named sulphur trioxide (SO_3); sulphur dioxide can also take, in the place of this additional proportion of oxygen, a quantity of chlorine, another compound being formed.

465. There are some elements which can combine both at ordinary and higher temperatures; different compounds being produced at the different temperatures: when this is the case, the lower combination is attended with little or no evolution of light, and the amount of heat developed is less than that produced by the higher combination. Example: Potassium combines at common temperatures with oxygen, forming potassium monoxide (K_2O); but when heated it burns with flame, and combines with four times as much oxygen. When phorus is exposed to the air, it always appears to emit a sh smoke, which in the dark is luminous. This effect is o a slow combination of the phosphorus with the oxygen e air, forming with it phosphorus trioxide (P_2O_3); if phorus be heated, it bursts, as we have seen, into a bril- flame, phosphorus pentoxide (P_2O_5) being produced, h contains two-thirds more oxygen than phosphorus ide.

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[1] One of the most convincing proofs that the air is a mixture of these gases is the manner in which the air behaves with water. If the air were a compound, the oxygen and nitrogen would be dissolved by the water in the same proportions as they exist in air, which is not the case. If, on the contrary, air is only a simple mixture of these two gases, and as they are not equally soluble in water, the composition of the dissolved gases in the water will be different from that of the air, which is the case. Another proof is, that if oxygen and nitrogen are mixed in the proportion constituting the air, we obtain a gaseous mixture identical with our atmosphere.

66. Having shown, by a number of examples and experi- s, that many elements can combine in more proportions one, it might be expected that the *theory* by which ists attempt to explain these and like facts would be d; but we think it better to here teach the student the *that many elements can combine in more proportions than giving rise to as many different compounds*, and explain the y in the next chapter.

67. We have compared in this chapter simple substances ements to letters, and compound substances to words. This ration may be carried still further; there are, as we all r, words having different functions, as noun, verb, &c., and ave now to learn that different chemical compounds possess rent functions.

68. The student will not have failed to notice in the ples and experiments that have been given, that the compds formed by the union of oxygen with other elements were d *oxides*; this is the general term for such compounds. many of the oxides of the metals have what are termed properties; they are therefore termed *basic oxides*. The es of the non-metals have properties the very opposite of e possessed by the basic oxides; these last are termed *acid s*. Here is the proper place for pointing out to the student emical distinction between a non-metal and a metal; the s of the non-metals never possess basic properties.

69. The oxides forming each of these classes combine water, forming compounds, the general term for which is *oxide*. The compounds formed by the union of the basic es with water are also termed bases. There is a very common ple of the conversion of a basic oxide into a hydroxide, h all my readers must have witnessed; it is the slaking of

lime. Lime (CaO) is a basic oxide, and by adding water to it there results the formation of the hydroxide (CaH_2O_2), in ordinary language called slaked lime. The combination of the water (H_2O) and the lime being attended with a great development of heat, a quantity of the water that has not combined is expelled in the form of steam. The compounds formed by the union of the acid oxides, or as they are also termed *acid-forming oxides*, with water are termed acids. Bases turn paper stained with turmeric *brown* in colour, and they turn *red* litmus paper *blue*. Acids do not affect the colour of turmeric paper, but they *redden blue* litmus paper.[1] Acids possess, as we know from the taste of vinegar (acetic acid), a sour taste. Bases and acids possess also the property of uniting with each other, a new class of compounds being formed, which are noticed in the next paragraph.

470. In ordinary language we call some substances *salts*; thus we speak of Epsom salts, common or table salt, saltpetre, and Glauber's salt ; these are salts, but there are many more known salts than these, for all the compounds formed by the union of acids and bases are termed salts ; and when the salt formed does not *change* the *colour* of *blue* or *red* litmus paper, the acid and base composing it are said to have *neutralised* each other.

471. There is another class of compound substances ; these compounds have neither acid nor basic properties, nor have they the composition of salts ; we distinguish them by the term *indifferent bodies.*

The student will have learned from what has been stated that compound substances can unite with each other.

Experiments on the combination of two compound substances.

359. Take two flasks ; place in one of them one ounce of a solution of ammonia, and in the other an equal quantity of hydrochloric acid ; fit a narrow straight tube to each of the flasks, by means of a perforated cork ; place them upon retort-stands, and adjust two bottles or flasks in such a way that the tubes may ascend into them; then boil the liquid in both vessels

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[1] The turmeric paper and blue and red litmus paper are called test papers ; they can be bought at a trifling cost at any operative chemist's. There are other test papers besides these, but these are sufficient for ordinary purposes. The teacher ought to show his class the action of acids and bases on these test papers.

for a few minutes. Ammonia and hydrochloric acid are gaseous bodies, and lighter than atmospheric air; consequently when their solutions are boiled they escape and collect in the bottles; if the mouths of the two vessels be brought together after they are filled with the gaseous compounds, the gases will combine and form a solid substance, a salt called ammonium chloride.

360. Add some dilute sulphuric acid to a dilute solution of ammonia, in such quantity that the solution, after the addition, has no action on blue or red litmus paper, which proves that the acid and base have united and formed a salt that has no action on test papers; the acid and base have therefore neutralised each other.[1]

472. Salts are capable of uniting with each other, thus forming more complex compounds.

473. We will give a few more experimental exercises, some of them illustrating what has already been stated, but in a somewhat different manner.

474. The more opposite substances are in properties the greater is their tendency to unite together; for example, acids and bases always unite when brought together, and acids dissolve the metals forming with them salts.

Experiments illustrating the action of Acids on Metals.

361. Add some nitric acid to some copper filings; the copper will be dissolved, the acid and the metal combining together; the solution of the compound, a copper salt, will be of a beautiful blue colour.

362. Add some dilute sulphuric acid to some iron filings; the iron wil. be dissolved, the acid and the metal combining together, and forming an iron salt.

475. These experiments are also examples that the compound formed by the union of two or more substances possesses properties perfectly different from those of its constituents. If, for example, the blue liquid which has been produced by the action of nitric acid on the copper were evaporated to dryness, we should obtain this blue crystalline substance[2]; or, if we evaporated the solution we have obtained by acting on iron with sulphuric acid, we should obtain this green-coloured salt. Now what relation or resemblance is there between the copper and the nitric acid and this blue crystalline solid; or between

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[1] What are termed standard solutions of sulphuric acid and ammonia can be purchased at the operative chemist's; on adding a given measure of each of the solutions together, the acid and the base will exactly neutralise each other.

[2] Specimens of cupric nitrate and ferrous sulphate ought to be on the lecture table.

this green crystalline solid and the sulphuric acid and iron? We learn from these examples that there is no resemblance between the properties of a compound and those of its constituents. *The properties of a compound cannot be predicted either from a knowledge of the properties of its constituents, or from any other source.*

476. As chemical affinity is a molecular force it acts only at *insensible* distances. For substances to combine, it is necessary, therefore, that they should be in intimate contact, or separated only by immeasurably small distances. On this account the solid state is less favourable to chemical action than the fluid state ; for in the solid state the particles of un like bodies cannot be brought into that close proximity which is essential for the development of their mutual affinities. The following exercises are given to teach the student the fact that solid bodies can rarely be made to combine by merely mixing them together, but if they are brought into a state of solution combination will at once take place if the substances capable of entering into combination.

Experiment illustrating that the solid state is less favourabl chemical action than the fluid state.

363. Mix together in a mortar some tartaric acid and sodium carbon each in the state of dry powder. No chemical action will take place long as they remain in the solid state, no matter how long they are together. After being well mixed, dissolve the powder in water instantly chemical action takes place (proved by the effervescence) beca in solution the substances are brought into more intimate contact.

477. Another peculiarity of the chemical force is that is exerted between different kinds of matter with diffe degrees of force. Nitric acid, for instance, will oxidise combine with most of the metals ; it oxidises and unites instance, with silver, mercury, copper, lead, and iron, b unites with these metals with very different degrees of inte with silver the combination is less powerful than with me with mercury less powerful than with copper ; with copp powerful than with lead ; and with lead less powerful tha iron.

Experiment on the relative intensity of affinity.

364. place a piece of metallic iron in a solution of cupric nit a quarter of an ounce of the copper salt in half a pint of water)

id will leave the copper, and combine with the iron, forming with it rrous nitrate; the metallic copper set free will be precipitated on the rtion of the metallic iron which remains uncombined.

Examination Questions.

365. Explain what is meant by chemical action, and give some illustrae examples.

366. What part does air take in the process of combustion?

367. If sulphur were heated in a vacuum, would it burn, or would it dergo any change?

368. When an element is consumed is there any gain or loss in weight?

369. Would any difference in the intensity of the light be obtained by pplying heated air to a gas-flame in place of cold air?

370. Why should the breath appear visible in cold weather and not at her times? From its appearing visible, what element or compound ould you infer is one of the substances in the respired air?

371. Two powders are given you: one is a mechanical mixture, the her a chemical combination of sulphur and iron. In what respects ould the two powders differ, and how could you prove that the mechanil mixture was not a chemical compound?

372. Give some proof that the air is a mixture, and not a chemical mpound.

373. Describe the properties of an acid.

374. Can more compounds than one be formed out of the same eleents?

375. Why will not carbon dioxide burn in air?

376. Describe the properties of a base.

377. What causes the blue flame which is frequently seen on the surce of coal-fires?

378. When are acids and bases said to have neutralised each other?

379. Why will not water burn in air?

380. Are the substances produced on burning a given weight of coal, ual, lighter, or heavier in weight than the coal itself?[1]

478. In giving a brief description of the properties of the ore commonly occurring elements, and the method for exacting them from their compounds, we will treat the two ivisions—non-metals and metals—apart; and we will comence with the first division.

479. *Hydrogen.*—This element is met with in nature in the ee state in small quantities—it is one of a mixture of gases in ertain volcanic emanations; it is given off from some of the il-wells; it has also been found occluded in meteoric iron; nd it is evolved during fermentation, and in the decay and ecomposition of organic substances. But it occurs almost ntirely in a state of combination; water is one of its most important compounds, and it is a constituent of nearly all animal and vegetable substances. It can be obtained, as has been shown (Experiment 400), by electrolysing water; this is one of the best methods for obtaining it in a pure state. We have learned that Na and K set it free from H_2O at ordinary temperatures: Li, Ba, Sr, and Ca likewise set it free from H_2O under the same conditions. Several of the other metals set it free at higher temperatures, the H_2O being passed over them in the state of steam, an anhydrous oxide of the metal being formed. Examples : Mg, Zn, Cd (monoxides of these metals being formed), and Fe, in the latter case a complex oxide, ferrosoferric oxide, Fe_3O_4, is formed. It is usually obtained by the action of H_2SO_4 or HCl on Zn, or on Fe (see note page 183), but there are numerous other metals, the alkaline, and alkaline earth, metals, also Al, Cd, &c., which displace the H in these acids. Zn, we have learned, can displace the H in KHO, or NaHO ; and Al sets free H from solutions of the alkalies, with probably the formation of a similar Al compound to that of the Zn one. With regard to its properties we have learned that it is the lightest body in nature. It combines with O at ordinary temperatures if the two elements are brought together in the *nascent* state ; and they combine together at the ordinary temperature when not in the nascent state by the action of the metal Pt, which by condensing them brings them sufficiently in contact for chemical union to take place ; this is called *surface* action, a kind of action the student has not yet been made acquainted with as regards gaseous combinations. Under no other conditions, as far as is known, does chemical union take place between H and O at ordinary temperatures; but at higher temperatures, we have learned, combination invariably takes place between them, and we have become acquainted with the phenomena which accompany the union.

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[1] *The rest of the chapter must be passed over for the present; the student will directed when to return to it.*

480. *Oxygen.*—We are already aware that this element is present in the free state in the air ; and we have also learned that it is the element which exists in the largest quantity, and is the most extensively diffused, in our planet. We have learned how to extract it from several of its compounds in Chapter XIV. ; and also in note page 174, the method usually adopted for obtaining it for lecture purposes ; but we will give

here two others, as we can by these methods obtain it from the atmosphere. (1) If air is passed over BaO when heated to a low red heat, it combines with O, forming BaO_2; at a bright red heat the latter compound is decomposed into BaO and O; this is an example that the temperature required for combination and that required for decomposition are not in some cases far apart. (2) If sodium or potassium manganate, Na_2MnO_4, is heated in a current of steam it parts with some of its oxygen, NaHO, and Mn_2O_3 being formed; after the O has been evolved the current of steam is discontinued and air is passed over it; the manganate is again formed. By this and the preceding process it will be seen O and N can be alternately obtained from the air. We learn from these two processes that O can be abstracted from the air and then given up over and over again by the same portion of substance. It has been combined, as we have learned, with all the elements but F; we have learned also that its combinations at high temperatures are not only attended with an evolution of heat but also of light; and that it is the gaseous body in the air that causes our fuel and light-producing substances to burn; it also causes metals to rust, and dead animal and vegetable matter to decay, by oxidising them. The respiration in animals is analogous to the burning of fuel; the oxygen inspired oxidises the *worn-out* animal matter, converting its carbon into CO_2 and its hydrogen into H_2O, which are expired; and the oxidation of these elements in the body, as their oxidation in the fire-place, is attended with an evolution of heat.

481. There is an allotropic modification of oxygen called *ozone*; it is one of the most powerful oxidising agents known, more powerful than ordinary oxygen; hence it oxidises bodies at ordinary temperatures, which only enter into union with ordinary oxygen at high temperatures. It is a colourless gas, with a peculiar odour, which has been compared to that of very dilute chlorine. Its molecular formula is O_3, three volumes of ordinary oxygen being condensed to two volumes in its formation. It is present in very minute quantities in the air, but it is only in the air of the country, and especially in that over the sea, that it is most constantly met with: in the air of towns it is almost always absent, owing to the quantity of oxidisable matter in the atmosphere. The usual method employed for detecting it in the air is by exposing paper moistened with a mixture of starch and KI; it sets free the I, which turns the starch blue; the following is its action upon KI:

$$2KI + O_3 + H_2O = 2KOH + I_2 + O_2.$$

This test requires to be employed with care as there are frequently present in the atmosphere other substances which colour the paper blue. Ordinary oxygen can be converted into ozone, but as yet only very small quantities have been so transformed. Ozone can be obtained in small quantities by one of the three following methods: '(1) It is evolved at the positive pole in the electrolysis of acidulated water. (2) It is obtained by the slow oxidation of phosphorus in the air. (3) It is formed by the discharge from an electrical machine through air or through oxygen.' It is not only converted into ordinary oxygen by heat, but even at ordinary temperatures it gradually changes into oxygen.

482. *Chlorine. Bromine. Iodine. Fluorine.*—These elements belong to one group. As it is as yet uncertain whether F has been obtained in the free state, we shall confine our remarks to the other three; not one of the three has been met with in nature in the free state. We are already acquainted with some of their physical, as well as their chemical, properties; we know that under ordinary conditions Cl is a greenish-yellow coloured gas, Br is a dark brown coloured liquid, and I is a solid of a blackish-grey colour and having a kind of metallic lustre. Common salt (NaCl) is the best known and most widely distributed compound of Cl. It may be obtained direct from NaCl as described in note, page 175; but it is more generally obtained by first obtaining HCl from NaCl (Ex. 644) then Cl from HCl (Ex. 670). I exists in combination with NaK, or Mg in small quantities in sea-water; it is also present in certain seaweeds, from the ash of which it is commercially extracted; it is also found along with Chili saltpetre ($NaNO_3$,); from the mother liquor of this salt it is also commercially extracted. The usual method adopted on the manufacturing scale is that given in Ex. 721. Br is present in small quantities in sea-water as a bromide: also in some sea-weeds from the ash of which it

is extracted by the same process as that adopted for the extraction of I. It is also present in the potash-beds of Stassfürth, also in certain mineral springs in America in quantities sufficient to make its extraction commercially profitable.

483. *Sulphur.*—This element is met with in large quantities in the free state in volcanic districts; but in a state of combination it occurs in much larger quantities, and is much more widely distributed. The two most extensive classes of sulphur compounds are the sulphides and the sulphates; the common ores of Cu, Pb, and Zn are sulphides, and there is a sulphur ore of iron, iron pyrites, FeS_2, which is extensively employed for converting, by roasting the ore, the S into SO_2 for the manufacture of sulphuric acid. The principal natural sulphates are those of Ca, Ba, Mg, Na, and Fe. Various sulphur compounds exist in the vegetable and animal world. Although it exists in large quantities in the free state in some countries, FeS_2 is submitted to destructive distillation for the purpose of obtaining it, thus—

$$3FeS_2 = Fe_3S_4 + S_2.$$

It is also obtained in the purification of coal-gas. It occurs in three forms in commerce, viz., roll sulphur or brimstone, flowers of sulphur, and milk of sulphur. The latter is obtained by decomposing a polysulphide of calcium with HCl; in this form it is in a very finely divided state, is of a white colour, and is used as a medicine. This element exists in several allotropic modifications.

484. *Phosphorus.*—This element does not exist in the free state of nature: it is very widely distributed, especially in combination with O and Ca, as calcium phosphate. This salt forms the chief inorganic constituent of the bones of animals, and it is from bones that P is obtained. The bones being freed from the organic matter and in the form of ash are treated with sufficient sulphuric acid to form an acid phosphate:

$$Ca_3(PO_4)_2 + 2H_2SO_4 = CaH_4(PO_4)_2 + 2CaSO_4.$$

The solution of the acid phosphate is poured off from the insoluble $CaSO_4$; the solution is evaporated to a syrupy consistence; it is then mixed with a sufficient amount of charcoal, evaporated to dryness, and afterwards gradually heated to full redness in an appropriate vessel: two-thirds of the phosphorus distils over, and is condensed by the water contained in the receiving vessel. The operation may be regarded as consisting of two stages:—1st, the conversion of the acid phosphate into calcium metaphosphate; 2nd, the setting free of the phosphorus:—

$$(1).\ \mathrm{CaH_4\,(PO_4)_2 = Ca2PO_3 + 2H_2O.}$$
$$(2).\ \mathrm{3(Ca2PO_3) + 10C = P_4 + Ca_3\,(PO_4)_2 + 10CO.}$$

We have learned that P unites with O at ordinary as well as high temperatures; but there is an allotropic modification of the element called red phosphorus from its colour, which does not combine with O at a temperature lower than 240° C. It is produced when ordinary phosphorus is heated, in an atmosphere of CO_2 or H, to a temperature of 240° C. for some time.

485. *Nitrogen.*—This is one of the two gaseous elements in the free state in the atmosphere, and it forms four-fifths of the bulk of the air; it also occurs in combination in ammonia, nitric acid, and other bodies, and is a constituent of many animal and vegetable substances. We have learned how it may be obtained from the air; it may also be obtained by heating a concentrated solution of ammonium nitrate. This element is distinguished by its inactive properties; it combines directly with but few of the elements, and it is difficult, we have learned, to make it combine even with these few; but it combines indirectly with a large number of the elements.

486. *Arsenic.*—The substance which popularly is called arsenic is not the element, as we have before stated, but As_2O_3; the element is of a steel-grey colour. It is widely distributed in nature; it occurs both in the free state and in combination with various metals. It is obtained by heating arsenical pyrites (Fe_2S_2As); the As is expelled and the S and Fe remain combined; it is also obtained by heating As_2O_3 with carbon. A small quantity of it is added to the lead in the manufacture of shot, as it causes that metal to assume more readily the spherical form and it also renders it harder.

487. *Carbon.*—This element is met with, as before pointed out, in three different modifications, viz., as the diamond, as graphite (black-lead), and in the form of charcoal, lamp-black, bone or ivory black, soot, &c. If we can consider some of the

elements more important than the rest, carbon is a most important one : *all* organic substances, whether animal or vegetable, are compounds of carbon; hence organic chemistry, or that department of chemistry treating of organic substances, has been defined as the chemistry of carbon compounds Although all organic substances contain carbon, all carbon compounds cannot be classed as organic substances. From what has been stated, the student will not be surprised on being told that the compounds of this element exceed those of any of the others both in number and variety. It only combines directly with a few of the elements. When any organic substance is heated in the absence, or in a deficiency, of air a portion of the carbon remains behind in its amorphous form. In all its three forms it is infusible at all temperatures ; and it is insoluble in any of the ordinary solvents ; the only solvent of it at present known is fused iron. With the halogens and with nitrogen it does not unite directly; its compounds with these elements can only be obtained by indirect methods. It does not combine directly with hydrogen, oxygen, or sulphur below 100° C. ; but at high temperatures its affinity for the two latter elements is very great ; it is on this account extensively employed in the extraction of metals from their ores, being, as it is termed, one of the most powerful reducing agents known.

488. *Boron.*—This element is not met with in nature in the free state ; and it is only met with in combination with oxygen, either as free boracic acid, or that acid in combination with bases forming the salts called borates. It has been obtained in two allotropic states, the amorphous and crystalline. It was until recently the only non-metallic element of which no hydrogen compound had been obtained, but recently this compound has been formed.

489. *Silicon.*—This element is not met with in nature in the free state; it is always met with in combination with oxygen, either in the state of silicic anhydride—as quartz, flint, &c.—or as silicic acid in combination with bases forming the extensive and important class of minerals termed silicates. It has been obtained both in the amorphous and crystalline state. It is the chief constituent, next to oxygen, of the solid earth.

490. *Metals.*—The elements belonging to this class vary

very much both in their physical and chemical properties; as examples we may notice the difference in their densities; K, Na, and Li are lighter, as we have learned, than water, whilst Pt is twenty-one times heavier than an equal bulk of that liquid; again, Au, Pt, and some of the rarer metals do not unite directly with O, whilst all the other metals do unite with it directly at ordinary or higher temperatures; K, Na, Sr, and Ba have such an affinity for it that they enter into union with it at common temperatures; and some of the metals, lead and iron for example, which in their *massive* state are unacted upon by air or pure oxygen unless moisture is present, when, in a very finely divided state, they come in contact with air instantly unite with oxygen, the combination being attended with ignition. *Alloys* are the intimate mechanical mixtures produced by melting metals together. Those alloys which contain mercury are called amalgams. Alloys and amalgams can only be produced from metals.

491. *Potassium. Sodium. Lithium.*—K and Na are obtained by placing a mixture of their dry carbonates and charcoal in an appropriate iron vessel and applying a strong heat: the metals distil over:—

$$K_2CO_3 + 2C = K_2 + 3CO.$$

These metals can also be obtained by the action of Fe on their hydrates at a strong white heat:—

$$4NaHO + 3Fe = 2K_2 + Fe_3O_4 + 2H_2.$$

Very recently another method has been invented for obtaining them from their hydrates, which is said to be so great an improvement that they can now be manufactured at as many pence a pound as formerly it cost shillings to obtain them. This method consists in heating their hydrates with carbide of iron (FeC_2), but in practice, instead of carbide of iron, an intimate mixture of Fe, in a very finely divided state, and pitch is used. Lithium cannot be obtained by the processes we have described for extracting the two previous metals; as yet it has only been obtained by the electrolysis of some of its salts. Owing to the great affinities of K and Na for O and Cl, they are employed for obtaining various non-metals and metals from their oxides and chlorides.

492. *Barium. Strontium. Calcium. Magnesium.*—These metals can be obtained either by the electrolysis of some of their salts, or by decomposing their chlorides with Na. The first three have no useful application ; but Mg is now manufactured on a large scale ; it is met with in commerce in the form of wire or ribbon. 'It may be ignited at the flame of a candle, and burns with an intensely brilliant white light, very rich in chemically active rays, a property which has led to its use in photography. Besides its application in photography, its light has been used in signalling ; it has been seen at sea at a distance of 28 miles.'

493. *Aluminium. Chromium.*—The first of these metals is obtained by passing the vapour of its chloride over heated K or Na. The preparation of the chloride from the mineral bauxite, a hydrate of alumina and ferric oxide, is somewhat complicated, and to describe it would occupy more space than we can afford. Al is used in the manufacture of various physical instruments ; it is also employed in the manufacture of its alloy with copper, which possesses the appearance of standard gold. If it could be obtained at a cheaper rate, it would have still more extensive applications, owing to its lightness, tenacity, unalterability in the air, &c. Chromium is obtained by heating its chloride with Zn in a crucible, and using a mixture of KCl and NaCl as a flux. The hardness of steel is greatly increased by the addition of 0·5 to 0·75 per cent. of this metal.

494. *Zinc. Cadmium.*—The first of these metals is extracted both from its carbonate and its sulphide ; the former ore is roasted to drive off the CO_2 and H_2O present ; the latter is roasted to expel the S in the form of SO_2 ; in both cases the ZnO is left in the state of oxide by the roasting. The ZnO is mixed with powdered coal, the mixture is placed in an appropriate furnace and heat applied ; the carbon of the coal removes the O ; the metal being volatile at a bright red heat distils over and is collected in appropriate receivers. Its most important uses are generally known. The latter metal is associated with Zn in its ores, and being more volatile than Zn distils over first, but, becoming in part oxidised, the distillate is mixed with small coal and distilled at a low temperature. It has no commercial importance.

495. *Iron.*—This metal, whose uses are so numerous, is, on that account, universally known, but, curious to say, the pure metal is not an article of commerce; very few indeed have seen it in a state of chemical purity; all the different varieties of iron met with in commerce contain carbon and some other foreign substances. All the substances associated with it exercise, no doubt, some influence on the character of the metal; but it may be said in general terms that it is chiefly to the amount and state of the carbon it contains that commercial iron in its different forms owes its characters; and to the varying amount and state of the carbon is due the several varieties of the metal known as pig or cast iron, wrought or malleable iron, and steel. It would be almost impossible to give, even in several pages, a very brief description of the methods adopted for the extraction of the metal from its various ores; we can therefore only glance at one or two points in the few lines we can devote to the subject. There are several varieties of pig iron; some contain more combined than uncombined carbon, in other kinds it is just the reverse. Pig iron contains a larger total of carbon than any other variety; steel contains the next largest quantity of C; and wrought iron contains the least. The metal is extracted from the native oxides and carbonates by first roasting the ores, and then smelting them in a furnace along with coal and limestone; the latter being employed to form a fusible mass with the gangue of the ore; this fusible mass is termed the *slag*.

496. *Manganese.*—This metal has no commercial importance; it may be obtained by subjecting a mixture of an oxide of the metal and charcoal to a strong heat in a crucible.

497. *Cobalt, Nickel.*—The first of these metals is not used in the metallic state in the arts; the latter is largely used in the form of rolled plates for nickel-plating; it is also one of the metals employed in the manufacture of the alloy known as German silver; and in some countries an alloy of nickel and copper is used as the material for the coinage in place of copper coinage. The two metals always occur associated together in nature, the one being in excess in some minerals, the other in other minerals; they are both constantly met with in metallic meteorites, and the student will probably have noticed that they have nearly the same atomic weight. Ni is prepared from the mineral NiAs and from an arsenical cobalt mineral called cobalt-speiss by removing the cobalt by a series of processes, and then reducing with charcoal.

498. *Copper.*—This metal can be very easily obtained from its carbonates, for they are reduced directly by heating with charcoal; but the ores of copper most generally smelted are sulphur ores; the chief one is a double sulphide of Cu and Fe ($Cu_2S.Fe_2S_3$) called copper pyrites; it requires about six different processes to remove the sulphur and iron from the copper; the smelting of this ore is therefore a very complex process. The iron pyrites now burnt for the production of sulphuric acid contains on an average three per cent. of copper, which all remains in the burnt ore; it is extracted from the burnt mass by roasting it with NaCl; the Cu is converted into a chloride, which dissolves out on treating the roasted mass with water, and is precipitated in the metallic state by placing scrap iron in the aqueous solution. This metal, which is one of those longest known to man, is employed for a great variety of purposes in the arts: it is very tenacious, ductile, and malleable. Brass is an alloy of copper and zinc: and bell-metal and gun-metal are alloys of copper and tin.

499. *Lead.*—This is one of the well-known metals; its most abundant ore is the sulphide, *galena*, PbS; and from this ore the metal is most generally extracted. The process most generally employed for extracting the metal from galena is by converting one-half the ore into sulphate by roasting; the whole quantity is then mixed together and the temperature raised, when the following reaction takes place.

$$PbSO_4 + PbS = 2Pb + 2SO_2$$

The metal can be worked very easily, and as it resists the action of air and moisture, it is employed for many purposes. Various alloys of the metal are used in the arts; *type-metal* is an alloy of lead, tin, and antimony; *plumber's solder* is an alloy of lead and tin. Galena almost invariably contains some silver as well as lead; the quantity is very variable, but even if it does not exceed three or four ounces in the ton of metal it is profitable to extract it, and the quality of the lead is improved by the removal. The process most generally employed is called Pattinson's process, from the name of the inventor. This process is founded on the fact that an alloy of lead and silver is more fusible than pure lead. The metal obtained on the reduction of the ore is melted in a large vessel, and then allowed to cool slowly; the portion which first crystallises is nearly pure lead; this is removed by a perforated ladle; the remelting both of the richer and poorer lead is continued until the largest quantity of the metal is all but free from silver, whilst the silver has increased in the richer portion to, say, 300 ounces of silver to the ton. This rich mass is placed in a *cupel* or shallow dish made of bone-earth (calcium phosphate), and again re-melted and exposed to a current of air, which oxidises the lead; the lead oxide thus formed melts and a portion flows off, whilst the rest is absorbed by the bone-earth; and the silver is left in the metallic state, and almost chemically pure; this process is called *cupellation*.

500. *Silver.*—This metal occurs in the metallic state, but it is in a state of combination that it occurs in the largest quantity. It is extracted from its ores by three different processes: 'either it is alloyed with lead and then the lead removed by oxidation, a process we have already described, or it is amalgamated with mercury and this removed by distillation; or, lastly, it is brought into solution as a salt and the metal precipitated by means of copper. The practical adoption of one or other of these processes depends upon the nature of the ore, the position of the mine, and the price of labour, fuel, &c.'

501. 'Pure silver is very little employed in the arts, as it is too soft to resist wear. In order to increase its hardness and tenacity, it is alloyed with a small proportion of copper, an addition which does not affect its colour, and in this form it is employed for plate, ornaments, coinage, &c. *Standard silver* is an alloy of silver and copper of a given composition fixed by law, and this standard varies in different countries. In England the standard contains 92·5 per cent. of silver.'

502. *Bismuth.*—This is not an abundant metal; it occurs in the metallic state disseminated through various rocks, from which it is separated by simple fusion; it is found also associated with nickel and cobalt, and from these ores it is also extracted. It is employed in the manufacture of fusible metals.

503. *Antimony.*—This metal is always extracted from antimonious sulphide, Sb_2S_3; this ore is first roasted, antimonious oxide and sulphurous anhydride being produced; the oxide is then reduced by carbon. It is only employed as a constituent of some alloys; it imparts to the alloy the property of expanding on solidification. The crude antimony of commerce is merely the sesquisulphide ore freed from the greater part of its earthy impurities.

504. *Tin.*—Independent of its value as a metal, tin is interesting to the inhabitants of Great Britain, as 'these isles became first known to the civilised nations of the ancient world by the commercial quest for tin, whether carried on by the direct voyages of the Phœnicians or Carthaginians (a question still in dispute), or by the overland route through Gaul, with the aid of its great rivers, to Massalia and other Greek colonies on the Mediterranean, which is fully described by Diodorus Siculus and Strabo. From the former source, Herodotus, in the fifth century B.C., may have derived his information about the islands in the extreme west of Europe, called Cassiterides, whence, he says, tin is brought to us.'[1] Tin is never met with in the metallic state; it is met with only in combination with oxygen, as SnO_2, called tin-stone. After the tin-stone has been freed from its earthy and metallic impurities, it is mixed with charcoal and reduced in a furnace. It is employed in coating other metals to prevent them from rusting, and it is a constituent of several alloys.

505. *Mercury.*—This metal occurs in its uncombined or native state, but it is chiefly met with in a state of combination; its most abundant ore is *cinnabar*, mercuric sulphide. It may be obtained from this ore either by burning off the sulphur and distilling the mercury; or by heating the ore with some substance which forms a non-volatile compound with sulphur, and distilling off the liberated mercury. It is used in silvering mirrors, and extracting gold and silver by the amalgamation process; it is of great use in physical and chemical laboratories

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[1] Smith and Groves's *Historical Atlas of Ancient Geography; Biblical and Classical.*

The alloys of the various metals with mercury are termed *amalgams.*

506. *Gold.*—This metal always occurs in its native state, and is always alloyed with silver, and sometimes with other metals. If it does not contain more than 10 per cent. of Ag it may be separated from that metal by melting the alloy in a clay crucible glazed inside with borax, and passing a current of Cl through the molten mass. The Ag is converted into chloride, which is prevented from volatilising by a layer of fused borax; if Zn, Sb, Bi, and Sn are present, they are also converted into chlorides, which are volatilised. The metal thus purified contains from 99·1 to 99·7 per cent. of Au. Its uses need not be stated, as they are so generally known.

507. *Platinum.*—This metal always occurs in its native state; it is mixed with palladium and other rare metals; these metals, in fact, are rarely found excepting associated with Pt. It is found only in comparatively few places, the largest amount being met with in Russia. On account of its extreme infusibility, its mode of extraction is complicated and peculiar. Owing to its being little acted upon by most chemical agents, and its high fusing-point, it is extensively employed in the manufacture of vessels for chemical purposes.

Examination Questions.

381. Calculate the weight and volume of oxygen, at the standard temperature and pressure, obtainable from 100 grams. of manganese dioxide,

a. by heat (Exercise 700).
b. by sulphuric acid (Exercise 661).

How would you extract silver from argentiferous galena?

382. If 100 grams. of steam is decomposed by passing it over red-hot iron, what will be the increase in weight of the iron, and what the volume in litres of the gas obtained measured at 0° C and 760 mm.?

383. From what you have learned as to the constitution of ozone, what is your inference, if any, as regards the constitution of allotropic bodies in general?

384. A splinter of a colourless body is given to you; how would you prove by a chemical test that it was diamond and not a piece of glass?

385. What are the materials you would require for making chlorine, and what quantities must be used in order to obtain 1,000 grams. of this gas?

386. Give the meaning of the terms gangue and slag.

387. What volume of air at 0° C. and 760 millimetres pressure is needed for the combustion of one gramme of hydrogen? What would be the volume of the product at 300° C.?

388. What are alloys, and what are amalgams?

389. Describe as fully as possible the means of obtaining hydrogen from water?

390. What are the sources of iodine; and how is it extracted from its compounds? Describe some of its physical and chemical properties, and state its vapour-density.

391. What is meant by saying certain substances are reducing agents?

392. Name the characteristic element in organic compounds.

393. For what purpose is the metal antimony employed?

394. Name the metals which, on being set free from their ores or compounds, are obtained by volatilising them.[1]

CHAPTER X.

LAWS OF COMBINATION—LAW OF DEFINITE PROPORTIONS—EXERCISES—LAW OF EQUIVALENT PROPORTIONS—LAW OF MULTIPLE PROPORTIONS—ATOMIC CONSTITUTION OF MATTER—GAY-LUSSAC'S LAW OF VOLUMES—EXPERIMENTS—EXERCISE—ATOMS AND MOLECULES—AVOGADRO'S LAW—TABLE OF THE MOLECULAR WEIGHTS AND VOLUMES OF THE GASEOUS AND VAPORISABLE ELEMENTS—THE VOLUME-WEIGHTS OF ELEMENTS AND COMPOUNDS—EXERCISES—ATOMICITY—GRAPHIC NOTATION—CLASSIFICATION OF THE ELEMENTS ACCORDING TO THEIR ATOMICITY—EXERCISES—QUESTIONS.

508. A chemical compound is distinguished from a mere mixture of substances, or from a simple adhesion of different substances, not only by its possessing properties entirely different from those of its constituents, but by the still more remarkable characteristic that it is always composed not only of the *same* elements, but they are always united in *unvarying* proportions *by weight.* We have to remember, then, that bodies always combine in *fixed* and *invariable* proportions by weight, and this law, it may be observed, is the basis of chemistry. No circumstances whatever can alter the proportional quantities of the constituents of a compound. Thus common kitchen or table salt—whether obtained from the salt-beds in Poland or from those in Cheshire, wherever found or however prepared, whether by the agency of man or by the operations of nature—is always made up of the two elementary bodies, chlorine and sodium, and in the invariable proportion of 23 parts by weight of sodium to

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[1] *The student must now pass on to par.* 601, *Chap. XI., and complete the study of that chapter.*

35·5 parts by weight of chlorine. Substances can, un
certain circumstances, assume different forms or states; f
instance, calcium carbonate exists both crystallised (marble) a
uncrystallised (chalk); water can assume the solid, the liq
and the gaseous states. Under all forms and states, howe
the chemical composition remains the same. Water—wh
in the solid state, as hail and snow; or in the liquid sta
rain or dew; or in the gaseous state, as steam,—is always m
up of the two elementary bodies, hydrogen and oxygen, a
always in the proportion of one part by weight of the fi
element to eight of the latter.

509. From these and similar facts we deduce the law, t
substances always combine in fixed and invariable proportio
weight, *which is not interfered with* when two or more substa
are brought together in proportions different from those in w
they will unite. If they combine they do so in the very s
proportions they would have done had they been bro
together in their natural proportions, *the only difference b*
that a portion—constituting the excess, of one or more of
substances—remains uncombined. Oxygen and zinc com
in the proportion of 8 parts by weight of the former to 32 p
by weight of the latter element; 40 parts by weight of the c
pound,—zinc oxide—being formed. Let us suppose that
are brought together in the proportion of 8 of oxygen to 3
zinc; the combining proportion of the two elements neve
less would be the same, 8 parts of oxygen combining wit
parts of zinc, 40 parts of oxide of zinc being formed, and 4 pa
zinc remaining uncombined. To take another assumed cas
Suppose they were brought together in the proportion of
oxygen to 32 of zinc, nevertheless 8 parts of oxygen, as be
would combine with 32 parts of zinc, 40 parts of oxide of
being formed and 1 part of oxygen remaining uncomb
Water and lime combine in the proportion of 9 parts by w
of the former compound, to 28 parts by weight of the l
37 parts by weight of hydrate of lime being formed. Sup
that we added 11 parts of water to 28 of lime, the consequ
would be that two parts of water would remain uncomb
37 parts of hydrate of lime being produced; or that we add
every 9 parts of water 31 parts of lime, 37 parts of h

lime would still be formed, whilst 3 parts of lime would remain uncombined.

510. This constancy of composition in chemical compounds, not only as regards their elements, but also as regards the weight of these elements, will be better understood, perhaps, if we select a few more examples, and arrange them in a different manner.

Water contains in 100 pts.		Hydrochloric acid contains in 100 pts.	
Oxygen	88·89	Chlorine	97·26
Hydrogen	11·11 × 8 = 88·88	Hydrogen	2·74 × 35·5 = 97·27
	100·00		100·00

Hydrosulphuric Acid contains in 100 pts.		Hydriodic Acid contains in 100 pts.	
Sulphur	94·12	Iodine	99·22
Hydrogen	5·88 × 16 = 94·08	Hydrogen	0·78 × 127 = 99·06
	100·00		100·00

511. It is evident from these examples that the different constituents of a chemical compound are present in different proportions by weight. Thus, in water, there is 8 times more, by weight, of oxygen than of hydrogen; for if we multiply the quantity of hydrogen contained in that compound by 8, the product will be the number representing the quantity of oxygen present in that liquid. In the same way, there is 16 times more by weight of sulphur than of hydrogen in hydrosulphuric acid; for if we multiply the quantity of hydrogen in 100 parts of that compound by 16, it will give us the number representing the quantity of sulphur the compound contains in 100 parts; and so with respect to the other two examples.

512. We may therefore express these facts in the following simple form. To one part by weight of hydrogen there is found—

In Water	In Hydrochloric Acid
8 parts of oxygen.	35·5 parts of chlorine.
In Hydrosulphuric Acid	In Hydriodic Acid
16 parts of sulphur.	127·1 parts of iodine.

Every 9 parts of water, then, are invariably composed of 8 parts of oxygen, and 1 of hydrogen; every 36·5 parts of hydrochloric acid are composed of 35·5 parts of chlorine, and 1 part of hydrogen &c.

513. *Chemical compounds are therefore always constant and* *definite in composition, not only as regards their components or elements, but likewise as regards the quantity by weight of these elements.* This is the first of the three great laws of chemical combination.

Exercises on the proportions by weight in which substances combine.

395. How much water would be produced from 28 lbs. of oxygen and 5 lbs. of hydrogen, and would either of the elements be in excess?

396. How much hydrochloric acid would be produced from 5 lbs. of hydrogen and 178 lbs. of chlorine, and would either of the elements be in excess?

397. How much hydriodic acid would be produced from 7 lbs. of hydrogen and 1,000 lbs. of iodine, and would either of the elements be in excess?

398. How much hydrosulphuric acid would be produced from 6 lbs. of hydrogen and 80 lbs. of sulphur, and would either of the elements be in excess?

514. The student became acquainted in the last chapter with two modes by which chemical compounds can be formed. (1) By *direct combination* of elements or compounds with each other; (2) By the *displacement* or *substitution* of one element, or compound substance, for another, see Experiment 363.[1] We must now give a few more examples of this latter method of forming chemical compounds, in order that the student may comprehend what is meant by **equivalent proportions.**

515. It has been shown that nitric acid dissolves some of the metals; it dissolves most of them. The chemical action is this: the metal removes from one portion of the acid sufficient oxygen to form with it a basic oxide; another portion of the acid oxide combines with this basic oxide, a salt being the result of the union; the acid oxidises and unites, for instance, with silver, mercury, copper, lead, and iron; but it unites with these metals with very different degrees of intensity. With silver the combination is less intense than with mercury; with mercury less intense than with copper; with copper less intense than with lead; and with lead less intense than with iron. If, therefore mercury were added to the compound produced by the combination of the acid with silver, the mercury would displace and take the place of the silver; if copper were added to the compound produced by the combination of the nitric acid and mercury, the copper would displace and take the place of the mercury; if lead were added to the compound produced by the combination of the nitric acid and copper the lead would displace and take the place of the copper; iron were added to the compound produced by the combination of nitric acid and lead, the iron would displace and take the place of the lead; therefore any metal in the list can displace all the metals which precede it, and can be displaced by each of the metals which follow it.

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[1] This experiment not only illustrates the fact that the solid state is less favourable to chemical action than the fluid state; but it also illustrates the fact that one compound substance can *displace* another from its combination; thus tartaric acid displaces the acid oxide, carbon dioxide, in the salt sodium carbonate; the carbon dioxide on being set free resumes its natural gaseous condition, and in bubbling up through the water causes the effervescence; by this displacement the original salt is, of course, destroyed, and a new one, sodium tartrate, is produced.

516. Mercury replaces silver in the proportion, by weight, of 100 parts of mercury for every 108 parts of silver; copper replaces mercury and silver in the proportion of 31·7 parts of copper for every 100 parts of mercury, or 108 parts of silver; lead replaces copper, mercury, or silver in the proportion of 103·7 parts of lead for every 31·7 parts of copper, or 100 parts of mercury, or 108 parts of silver; iron replaces lead, copper, mercury, or silver, in the proportion of 28 parts of iron for every 103·7 parts of lead, or every 31·7 parts of copper, or every 100 parts of mercury, or every 108 parts of silver. We will illustrate the proportions in which substances replace and are replaced by a few more examples:—Sixteen parts of sulphur, we have learned, combine with 1 part of hydrogen and form 17 parts of hydrosulphuric acid; if we were to replace the sulphur in hydrosulphuric acid by iodine we should require 127·1 parts of iodine for every 16 of sulphur, and the compound, after the sulphur had been replaced by the iodine, would not be hydrosulphuric acid; it would be hydriodic acid: if we were to replace the sulphur in hydrosulphuric acid, or the iodine in hydriodic acid, by chlorine we should require 35·5 parts of chlorine to replace every 16 parts of sulphur, or every 127·1 parts of iodine. From these examples it will be seen that whether a substance, A, combines with another substance, B, by direct combination, or by replacing some other substance with which B is united, it (A) always combines with B in the same proportions by weight in whichever way the combination takes place: thus, chlorine always unites with hydrogen in the proportion by weight of 35·5 parts of chlorine to every 1 part of hydrogen, whether the chlorine unites with the hydrogen by direct combination, or by displacing some substance with which the hydrogen was united.

517. From these and like facts the second of the laws of chemical combination, and which is called the *Law of Equivalent Proportions* has been deduced, viz., that *when one element or compound takes the place of another element or compound, the replacement takes place in different but definite amounts.*

518. The third law is called the *Law of Multiple Proportions*; it has been deduced from innumerable facts the chemist has discovered ; he has found, as we have learned, that some substances combine in more proportions than one, forming totally distinct compounds ; and that the higher proportions are multiples of the lowest combining proportion of the substance : but before we can enter more fully with advantage to the student on this subject, it is desirable he should be made acquainted with certain theories.

519. It was by the aid of the balance that chemists discovered that substances unite in certain definite proportions by weight ; and by the aid of the balance they have determined the proportions by weight of the different substances that unite ; and the proportions by weight in which they replace each other, in other words their *equivalent proportions.* Scientific facts are truths, and are consequently unalterable ; bnt the explanations or interpretations scientific men give of those facts are at the best only guesses at truth. Different theories may, and often are, devised by different scientific men to explain the same series of facts ; and even the theory which is most generally accepted is liable to change with the progress of the science, as we shall learn from the theory we have now to explain.

520. The cause of elements and compounds always uniting in unvarying proportions by weight is unknown to us ; all that we know from chemical investigations is that it is a principle impressed on matter by the Creator ; and we further see how important it is ; for if substances did not unite in unalterable proportions by weight there would be no stability in the composition or properties of natural and manufactured chemical substances : what at one time might be a food, might at another time be a poison. Although we do not know why substances always unite together in unvarying proportions by weight, a beautiful theory was invented by the late Dr. Dalton to account for the fact; it is known as *Dalton's Atomic Theory*, and although purely speculative, has exerted an important influence in aiding the progress of chemical science. The ancient philosophers had conceived, before it had been discovered that the elements unite in certain fixed and definite proportions by weight, that matter was made up of small *indivisible* particles they termed *atoms*. Dalton adopted this view of the constitution of matter, but he conceived that the atoms of the different elements differed in weight, which the ancients had not conceived, for they had no knowledge of elementary substances; *fire*, *earth*, *air*, and *water* they regarded as the elements. It will be seen that, if Dalton's theory be admitted, the difficulty of explaining why the elements should always combine, and replace each other, in certain *definite* proportions by weight is removed; for according to this view, it is one or more atoms of one element combining with one or more atoms of another element: for *atoms*, unlike *molecules*, are *chemically*, as well as *mechanically*, *indivisible*. The absolute weight of the atoms cannot of course be determined, because they cannot be *isolated* or got into a separate state; the atomic weights are therefore *relative*, not *absolute* weights.

521. Hydrogen being the lightest substance that is known, its combining proportion by weight is less, as we have seen, than that of the other elements; Dalton, therefore, assigned as the weight of its atom the least whole number, viz. one; and as at the time he propounded his Atomic Theory there was only one compound of hydrogen and oxygen known, viz. water; and as water is composed of these two elements in the proportion, as we have already shown, of one part by weight of hydrogen to eight parts by weight of oxygen, Dalton considered that in the formation of water one atom of each of the elements unite together; therefore, if the atom of hydrogen be taken to weigh *one*, the atom of oxygen, according to this interpretation, must weigh *eight*.

522. But Dalton in framing his system of Atomic Weights neglected to observe, or at all events to take into account, the proportions by volume in which gaseous substances combine. This oversight is not surprising if we notice the few elements that are naturally gaseous, and the comparatively few that can be rendered gaseous ; and it had been by *weighing* that the discovery had been made that substances unite and replace each other in certain definite proportions by weight. About the time Dalton was constructing his Atomic Theory, Humboldt and Gay Lussac observed that oxygen and hydrogen unite by *volume*, to form water, in the proportion of 1 volume of the former to 2 volumes of the latter element ; therefore, if Dalton's hypothesis were correct, the hydrogen atom must occupy double the volume the oxygen atom occupies. Now, if this were the case, we should expect a difference in their dilatation and compression when subjected to the same variations of temperature and pressure ; but it has been proved by experiment that there is no difference in this respect between the two gases ; they expand and contract alike, the pressure being the same for equal additions or subtractions of heat, and they also experience the same change in volume for equal pressures. This uniformity in expansion and contraction extends to all gases and also to vapours, at some distance above their points of condensation. This uniformity of gases in their relations to heat and pressure led scientific men some years ago to believe that in the same *volume* all gases contain the same number of ponderable atoms set at equal distances, and whose natural repulsion is expressed by the same law.

523. But this generalisation—that equal volumes of gases under the same conditions of pressure and temperature contain an equal number of atoms—is true only within certain limits. In the sense which is now attached by chemists to the word atom the proposition is true only for a certain number of the gaseous and vaporisable elements : hydrogen, oxygen, nitrogen, chlorine, bromine, for example ; but it is not correct as regards phosphorus, arsenic, mercury, &c., as we shall learn as we proceed, for they do not contain in equal volumes the same number of atoms as hydrogen, oxygen, chlorine, &c. Neither is it correct, as we shall learn, as regards compound gases ; although they are subject to the same laws of expansion as the simple gases. What is accepted is that *equal volumes of gases contain the same number of molecules*, and that a molecule of each substance occupies two volumes if one atom of hydrogen occupies one volume. Equal volumes, however, of hydrogen and oxygen contain the same number of atoms; and as one volume of oxygen is sixteen times heavier than one volume of hydrogen, and as equal volumes of the two gases contain the same number of atoms; it follows that the atom of oxygen must be sixteen times heavier than the atom of hydrogen, and not eight times, as Dalton thought; and therefore, the correct formula for water will not be HO, but must be H_2O, as two volumes of hydrogen unite with one volume of oxygen to form one molecule of water which occupies two volumes.

524. Although in everyday life we all see liquids measured as well as solids weighed, yet I have, and I believe other chemical teachers have likewise, found that students less readily comprehend volume-relations than they do those relating to weight; we will therefore, before going further into the subject, give a few experimental exercises, showing the volumetric proportions in which the elements hydrogen, chlorine, oxygen, and nitrogen combine to form the compounds hydrochloric acid, water, and ammonia; and the proportion by volume in which hydrogen combines with the non-volatile element carbon to form the gaseous compound *methane*, as the student will then learn by observation the principles of volume-combination.

Experiments illustrating the volumetric constitution and condensation ratios of gaseous compounds.

399. The following experiment proves that hydrochloric acid can be formed by the direct union of its elements, hydrogen and chlorine; and that they unite in *equal* volumes, and that the volume the hydrochloric acid occupies is the *sum* of the volume of its constituents, there being *no condensation.*

'Two stoppered glass vessels of exactly equal capacity (fig. 66) are united by a tube, which is closed by a stopcock. After closing the stopcock, one of these is filled with chlorine and the other with hydrogen, and the stoppers are replaced. On opening the stopcock in the dark, diffusion will cause the gases to mix, an action which will be more rapid if the part of the apparatus containing the chlorine be placed uppermost. If the apparatus be now exposed to the diffused light of a well-lighted room, but not to direct sunlight—otherwise an explosion will occur—the colour of the

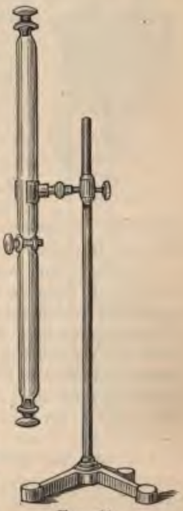


FIG. 66.

the limb with the vapour, we have to enclose it in a high glass cylinder *d* (fig. 70), the lower mouth of which is fastened around it by means of a perforated cork, whilst its upper mouth (likewise closed by a cork) rises about five centimetres above the sealed extremity. The glass tube *e* is connected with a tube from the flask containing the amylic alcohol; the vapour of the alcohol, after protracted ebullition of the liquid, passes into the annular space and raises the temperature of the gases in the sealed limb to 132°; this increase of temperature causes them to expand. The mercury must be kept at a *uniform* level in both limbs, either by adding or withdrawing a portion of the metal, the latter by the stopcock *b*. When the gaseous mixture ceases to expand, the height is marked by slipping a caoutchouc ring over the outer glass cylinder; a little more mercury, say an inch, is then poured into the outer limb of the U-tube, which is then tightly closed by a cork. A spark is now passed from the induction coil [this part of the apparatus is not shown in the fig.] by the wires − and +. After the explosion the cork is removed from the open limb, and the mercury levelled in the two limbs, when the volume of steam will be found to be just two-thirds of the volume of the gaseous mixture before the explosion, if the experiment has been correctly made. On cooling down, the steam condenses, and the mercury, with the exception of the small space the water occupies, entirely fills the closed limb of the U-tube.

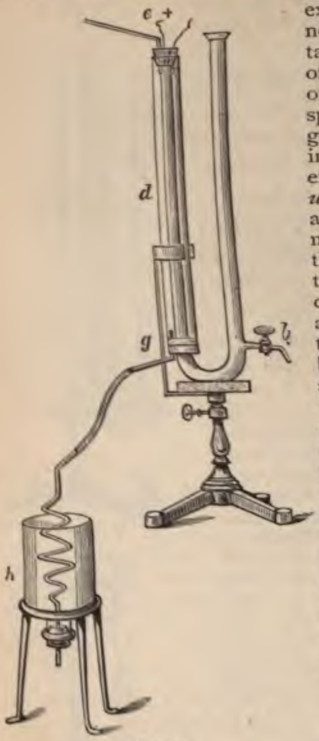


FIG. 70.

To prevent the alcohol vapour from escaping into the atmosphere, the lower extremity of the glass cylinder is connected, by means of the glass tube *g*, with the condenser *h*; the worm of the condenser being surrounded with cold water. The tube conveying the alcohol vapour into the glass cylinder is jacketed with caoutchouc to prevent loss of heat.

527. By the next experiment we learn the proportions by volume in which nitrogen and hydrogen combine to form the gaseous compound ammonia (NH_3), and the amount of condensation that takes place on their union. The compound is best prepared by mixing one part of ammonium chloride in powder with two parts of slaked lime, and heating the mixed mass in a flask to which a bent tube is connected, and with that a drying tube containing unslaked lime; to the other end of the drying tube a bent tube is attached, to convey the gaseous compound into the eudiometer.

402. 100 cubic centimetres of ammonia are passed into a eudiometer tube (fig. 47). Electric sparks, furnished by an induction coil, are passed through the gas by means of the platinum wires as long as any increase in the volume of the gaseous mixture takes place, as the ammonia is decomposed by the electric spark. After allowing the gases to cool, the volume is measured, and it is found that the gaseous mixture just measures double the volume the gaseous ammonia measured. We have learned, then, that two volumes of ammonia furnish, on decomposition, four volumes of its constituents; but we have yet to learn what is the volume-proportion of the two gases. We add for this purpose to the 200 c.c. of the mixed gases 100 c.c. of oxygen. We then explode by the electric spark, and only the hydrogen under these conditions unites with the oxygen; after allowing time for the gases to return to the temperature of the air, we observe the volume, and we find that it only measures 75 c.c.; there has, therefore, been a contraction of 225 c.c.; two-thirds of this, viz. 150 c.c., represents the amount of hydrogen in the 200 c.c. of the mixed gases; consequently, there was present 50 c.c. of nitrogen in the 200 c.c. The two gases, therefore, combine to form ammonia in the proportion of three volumes of hydrogen to one of nitrogen; and on their union there is a condensation to two volumes.

528. By the next experiment we shall learn the volume-proportion of one of the constituents of the compound methane, or marsh-gas as it is frequently called, because it is produced in the decomposition of vegetable matter under water, and is given off when the mud at the bottom of stagnant pools is stirred: it also constitutes the 'fire-damp' of coal mines. Methane, CH_4, is a compound, as shown by the symbol, of carbon (one of the elements which has never been volatilised, although this and many of its other compounds are gaseous bodies), and of hydrogen. This hydrogen compound differs from the three preceding ones, viz., HCl, H_2O, 'and H_3N, that whereas both constituents of these are gaseous, in this fourth one the hydrogen is united with a body which is not only solid at common temperatures, but incapable of volatilisation by any, the intensest, means at our present command. Hence, while the first three compounds illustrated, with an admirable symmetry, and in an unbroken ascending scale, the laws of combination and progressive condensation, by volume as well as by weight, the fourth exemplifies volume-combination and condensation only so far as its gaseous ingredient is concerned; our *positive* knowledge, as to its solid constituent extending only to the *weight*-ratio, and any views we may entertain as to the *volume*-ratio of this element being of necessity speculative. It is indeed, maintained by many chemists, and not without some show of reason, that analogy affords a fair basis for such speculation ; but not until carbon shall have been actually volatilised, and its vapour actually weighed, shall we be entitled to rank marsh-gas, with respect to its volumetric constitution, on the same certain footing as experiment already assigns to hydrochloric acid, water, and ammonia.'

403. To prepare marsh gas, mix very intimately one part of sodium acetate with four parts of soda lime ; both chemicals can be obtained of the operative chemist. Place the mixture in a flask of glass (or better of copper or iron), arranged for gas-disengagement ; then heat the flask to the point at which the gas begins to be evolved. It is eight times heavier than hydrogen ; hence, taking the unit-volume of hydrogen as one, two volumes of the gas weigh sixteen.

The experiment for ascertaining how much hydrogen is present in two volumes of the gas is performed in the U-tube, fig. 70 ; it is decomposed by causing a spark current of the induction coil to traverse the wires in the tube. The decomposition proceeds rapidly at first, but afterwards more slowly, so that a considerable time is required to bring the operation to a close. The volume of hydrogen, when the experiment is complete, ought to be double the volume of the marsh-gas operated upon, but on allowing the mercury to run out at the tap till it has become level in the two limbs of the tube it will not be found, even in the most successful experiment, to be quite double the volume, owing to a minute portion of the methane escaping decomposition. The experiment teaches us that there are four volumes of hydrogen in two volumes of methane, and two volumes of this carbon compound weigh sixteen, and as four volumes of hydrogen weigh four, the amount of carbon by weight in two volumes is, of course, twelve.

529. We will give one exercise, and in this one the student will have to explain each step in the experiment after the teacher has made it ; and he must further state the proportions by volume in which the two elements combine, and what condensation, if any, takes place on their union.

404. The gaseous compound selected is one of the oxides of nitrogen ; from the effects it produces when respired in small quantities it is known under the name of laughing gas. It is prepared by heating, gently at first, ammonium nitrate, previously dried in a flask to which a delivery tube is attached, and collecting the gas over mercury or warm water.

100 c.c. of the gas are collected in the eudiometer tube, fig. 47 ; 150 c.c. of hydrogen are then added, and the mixture is exploded by the electric spark. After the gaseous substance which remains has cooled down to the surrounding temperature, it is found to measure 150 c.c. ; 100 c.c. has therefore disappeared. We are unable as yet to deduce the composition of the oxide of nitrogen gas ; and the student must state the

	A Iron	B Oxyg
Monoxide of iron	56	16
Sesquioxide of iron	56	24
Triferro-tetroxide	56	21

564. The monoxide is not strictly an example of this it is introduced because it is the lowest oxide of iro observe that its molecule is composed of one atom of the and one atom of the non-metal; it is therefore represent the formula FeO. The oxygen in the other two o increases, we see, on the quantity in the monoxide quantities representing 1½ atoms of oxygen in the sesquiox and 1⅓ atoms in the triferro-tetroxide. Now, as we have alre stated, quarter or half atoms are inadmissible, being inconsist with the supposition on which the atomic theory is based; the and like compounds must, therefore, be composed of 2 or mo atoms of A united with 3 or more atoms of B. In the sesqu oxide 2 atoms of iron are united with 3 atoms of oxygen; i the triferro-tetroxide 3 atoms of iron are united with 4 atoms c oxygen.

565. Representing, as we have done in the previous series, the constituents of a binary compound (a compound composed of two elements) by A and B, the following illustrates this series:—

$$2A+3B, 2A+4B, 2A+5B, 3A+4B, \text{\&c.}$$

566. The method for finding the rational proportion of substances united in complex proportions must now be explained to the student. An atom, as we have stated before, is considered to be a particle of matter so small that it is incapable of division; therefore, when we find, by analysis, the elements of a compound united in the proportion by weight of 1 atom of A to 1½, or 2½, or 3½ atoms of B, as these proportions are inconsistent with the supposition on which the atomic theory is based, we get rid of the fractions by doubling both numbers; thus, 2 atoms of A to 3, or 5, or 7 atoms of B: the proportions, by weight of the elements, still remain the same, whilst the dividing of an atom is avoided. When the elements are found to be in the proportion of 1 atom of A to 1¼ atoms of B, this is in the proportion o 4 atoms of A to 5 atoms of B; if, therefore, we multiply the 1 atom of A and 1¼ atoms of B by 4, we obtain

molecules and atoms. We have already stated that a molecule is the smallest mass of matter that is *mechanically indivisible*, but *chemically divisible* ; an atom is the smallest mass, for it is both *chemically* as well as *mechanically indivisible* ; there are therefore, the student must remember, no half-atoms. No smaller mass of matter than an atom *can exist*, and it *only exists* in a chemical compound, not in a *free* state ; a molecule being the smallest mass that *exists* in the *free state*.

534. We have learned that the same variations of temperature and pressure produce in all gases the same changes in volume ; from this Avogadro in 1811 deduced the law that *equal volumes of all gases at the same temperature and pressure contain an equal number of molecules* ; this law is now one of the chief foundations of modern chemistry.

535. Molecules, as well as atoms, are *invisible*. We gave, in par. 9, Sir W. Thomson's illustration of the size of a molecule. Although a molecule of any substance is invisible, we can ascertain the *relative* weight of a molecule of any gaseous or vaporisable substance, as we shall show presently ; and we can *chemically divide* it, by decomposing it into its constitutents.

536. That the molecule of a compound, although mechanically indivisible, is a *larger mass* of matter than the portions of matter of the different elements of which it is composed will be obvious to all my student-readers ; these smaller portions of matter we term *atoms*. The word atom is derived from two Greek words signifying a particle of matter so minute as to admit of *no* division.

537. That the molecules of compounds composed of *dissimilar* elements are formed out of a combination of the atoms of the elements composing them is self-evident ; but the question naturally arises, Are the *atoms* of the elements themselves in the free state *united*, or do they *exist singly* ; in other words, is a mass of any element an agglomeration of the molecules of that element or of its atoms? In answering this question, we will commence by informing the student that it is only of recent date that chemists have called the smallest mass of an element in the free state a molecule. It was formerly termed an atom, as well as the smallest mass which is chemically indivisible. It was also formerly considered that only dissimilar elements could unite together ; now, it is admitted that the atoms of the same element can chemically unite with each other ; and that the homogeneous atoms, in some cases, are united by a stronger affinity than are the heterogeneous atoms of some compounds. 'For example, the molecule of nitrogen is a much more stable compound of the two atoms of the element than the compounds of nitrogen with oxygen, or with hydrogen or chlorine.' We shall adduce chemical evidence in proof that the atoms of the elements are, as a rule, united together in the free state in the last chapter ; but here we will call the attention of the student to the fact that the molecules of the same substance, whether composed of dissimilar or similar elements, never exercise chemical action upon one another.

538. It is now believed, as has been implied, that the elements are, as a rule, united in pairs ; thus the molecule of hydrogen is composed of two atoms of that element ; it is therefore symbolically represented thus, HH or H_2, whilst the atom is represented by only one H : these symbols not only signify the molecule and atom, but they also *represent* the weight of the molecule and atom of the element, the molecular weight being double the atomic weight ; and we take, as we have already learned, the hydrogen atom as the *unit* of the atomic weights of the elements.

539. The molecule of hydrogen, of oxygen, of chlorine, of nitrogen, and of some of the other elements, as is shown in the following table, is composed of two atoms of the element ; the symbols of their molecules are therefore represented as follows. HH or H_2, OO or O_2, ClCl or Cl_2, NN or N_2, and of course the molecule of these elements occupies *double the volume* the atom occupies. And the *specific gravity* of their molecules is double the sp. gr. of their atoms ; when the molecular weight of the element is twice that of its atomic weight, the molecule is *diatomic*, that is composed of two atoms ; and in a notation devised by Dr. Hofmann one volume of the element is represented by a *square*, and the molecule by a *double* square, as shown in the following table.

540. The number of atoms in a molecule of an elementary body can be found by dividing the molecular weight, as deduced carbon, thus C'''' or C^{IV}, &c. The Roman numerals are more convenient than the dashes for the higher numbers, and are, therefore, more frequently employed. It is not necessary to indicate the atomicity of a monad element, for if no accent or numeral be attached to the symbol of an element it is understood that that element is a monad.

571. From the illustrations we have just given of the atomicity of the elements, the student may have very naturally concluded that each element has *only one* atomicity ; but when we recall to his recollection what he is already acquainted with, viz., that there are at least two oxides of carbon, and five oxides of nitrogen, he will at once see that carbon and nitrogen must have more atomicities than one ; for if they had only one, there could only be one oxide of carbon, and one oxide of nitrogen. The *varying* of the atomicity of the same element is explained by supposing that sometimes some of the points of attraction are not *active*, but are *latent* or in a *state* of *rest* ; thus, when the points of attraction in nitrogen are all active it is a pentad, as in the salt ammonium chloride (NH_4Cl), for as the formula shows, one atom of nitrogen in that compound is combined with five atoms of monad elements ; whereas in ammonia (NH_3) it acts as a triad, two of the points of attraction being, it is imagined, in a state of rest or latent. When an element has entered into combination with another in its *greatest* atomicity, as nitrogen in NH_4Cl, the compound is said to be *saturated*, for it is incapable of entering into direct combination with another element or compound ; but when the combining units are not all united the compound is said to be *unsaturated*, because it can enter into further combination with an element or compound ; thus ammonia (NH_3) combines with HCl, forming the saturated compound NH_4Cl. The variation in the atomicity of an element takes place, with very few exceptions, by an *even* number of combining units ; thus carbon is either a *tetrad* or a *dyad* ; sulphur is either a *dyad*, a *tetrad*, or a *hexad*.

572. Two elements having the same atomicity unite together in equal atomic proportions ; but, with a few exceptions, elements of different atomicities do not unite in equal atomic proportions. A dyad and a triad unite in the proportion of three atoms of the dyad to two atoms of the triad. When a tetrad and a dyad unite together, they unite in the proportion of one atom of the tetrad to two atoms of the dyad; a tetrad and a triad unite in the proportion of three atoms of the tetrad to four atoms of the triad. A pentad unites with a dyad in the proportion of two atoms of the pentad to five atoms of the dyad; with a triad in the proportion of three of the pentad to five of the triad; with a tetrad in the proportion of four of the pentad to five of the tetrad. A hexad unites with a dyad in the proportion of one of the hexad to three of the dyad; with a triad in the proportion of one of the hexad to two of the triad; with a tetrad in the proportion of two of the hexad to three of the tetrad; with a pentad in the proportion of five of the hexad to six of the pentad. When elements replace each other in compounds, they do so in these proportions: a monad replaces another monad in equal atomic proportions. This is also the case with other elements of equal atomicities; but one atom of a dyad replaces two atoms of a monad, a triad replaces three atoms of a monad, a tetrad four atoms, a pentad five atoms, a hexad six atoms; a triad replaces a dyad in the proportion of two atoms of the triad to three atoms of the dyad; in short, the elements replace each other in the atomic proportions in which they combine with each other.

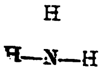
573. Elements with an even number of atomicities are termed artiads, and those with an uneven number are termed perissads.

574. Each unit of atom-fixing power is frequently termed a bond; and a graphic notation has been introduced into chemical works of late years, in which the elements of a compound are connected together by a number of lines or bonds corresponding with their degree of atom-fixing power. We will represent the four typical hydrogen compounds and ammonium chloride by this kind of notation to illustrate what has just been stated:—

Hydrochloric acid, HCl . . . . $H—Cl$

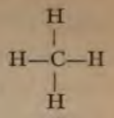
Water, H_2O . . . . $H—O—H$

Ammonia, H_3N . . .



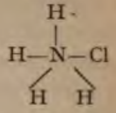
$$\begin{array}{c} H \\ | \\ H—N—H \end{array}$$

Marsh gas, H_4C . . . . . .



$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

Ammonium chloride . . . . . .



575. In the following table the more commonly occurring elements are grouped according to their atomicities, and the atomic weights of these elements are given. We have given those atomicities of the elements which the student requires to know first, according to the system of teaching we have adopted in this book ; as he is made acquainted with compounds in which some of the elements composing them display other atomicities, his attention will be directed to it, and the necessary information given. He will observe that the termination of the name of the element varies in the two atomicities, *ous* being the termination of the lowest atomicity, and *ic* of the higher atomicity. Example : iron in the ferr*ous* state, iron in the ferr*ic* state. The student must again refresh his memory with the names and symbols of the elements ; this is all-important, for he can make no satisfactory progress in chemical nomenclature and notation unless he has committed the names and symbols perfectly to memory, and he should also make himself perfectly acquainted with the atomicities of these elements as given in the table.

Exercises on Combination by Weight.

426. Every 100 parts of a gold chloride are composed of 64·79 of Au and 35·21 of Cl. How much Cl is united with 196 (weight of one atom) of gold, and how many atoms of Cl is it equal to, and give the symbol of the chloride ?

427. Every 100 parts of an iodide of gold are composed of 60·63 of Au and 39·32 of I. How much I is united with 196 of Au, and what is the symbol of the iodide ? and what is the atomicity of the gold in this compound and the one in the preceding exercise ?

428. Every 100 parts of alumina, an oxide of aluminium, are composed of 53·4 of Al, and 46·6 of O. How much O is united with 27·5 (weight of one atom) of Al ? and what symbol or formula would you assign to the compound ?

429. Every 100 parts of a boron chloride are composed of 9·36 of B and 90·64 of Cl. How much Cl is united with 35·5 (weight of one atom) of B, and what is the symbol of the compound ?

TABLE.

Classification of the more commonly occurring Elements according to their atomicity.

Names of the Elements	Symbols	Atomic Weights
MONADS—		
Hydrogen	H	1
Chlorine	Cl	35·5
Bromine	Br	80
Iodine	I	127
Fluorine	F	19
Lithium	Si	17
Potassium	K	39
Sodium	Na	23
Silver	Ag	108
Mercury in the mercur*ous* state	Hg	200
Copper in the cupr*ous* state	Cu	63
DYADS—		
Oxygen	O^{II}	16
Barium	$\mathrm{Ba}^{\mathrm{II}}$	137
Strontium	$\mathrm{Sr}^{\mathrm{II}}$	87·5
Calcium	$\mathrm{Ca}^{\mathrm{II}}$	40
Magnesium	$\mathrm{Mg}^{\mathrm{II}}$	24
Zinc	$\mathrm{Zn}^{\mathrm{II}}$	65
Cadmium	$\mathrm{Cd}^{\mathrm{II}}$	112
Cobalt	$\mathrm{Co}^{\mathrm{II}}$	58·8
Nickel	$\mathrm{Ni}^{\mathrm{II}}$	58·8
Lead	$\mathrm{Pb}^{\mathrm{II}}$	206·4
Sulphur in the sulphur*ous* state	S^{II}	32
Mercury in the mercur*ic* state	$\mathrm{Hg}^{\mathrm{II}}$	200
Copper in the cupr*ic* state	$\mathrm{Cu}^{\mathrm{II}}$	63
Iron in the ferr*ous* state	$\mathrm{Fe}^{\mathrm{II}}$	56
Manganese in the mangan*ous* state	$\mathrm{Mn}^{\mathrm{II}}$	55
Tin in the stann*ous* state	$\mathrm{Sn}^{\mathrm{II}}$	118
TRIADS—		
Boron	$\mathrm{B}^{\mathrm{III}}$	11
Gold	$\mathrm{Au}^{\mathrm{III}}$	196
Aluminium	$\mathrm{Al}^{\mathrm{III}}$	27·5
Chromium	$\mathrm{Cr}^{\mathrm{III}}$	52·5
Phosphorus	$\mathrm{P}^{\mathrm{III}}$	31
Nitrogen	$\mathrm{N}^{\mathrm{III}}$	14
Iron in the ferr*ic* state	$\mathrm{Fe}^{\mathrm{III}}$	56
Arsenic in the arseni*ous* state	$\mathrm{As}^{\mathrm{III}}$	75
Antimony in the antimoni*ous* state	$\mathrm{Sb}^{\mathrm{III}}$	122
Bismuth in the bismuth*ous* state	$\mathrm{Bi}^{\mathrm{III}}$	210
TETRADS—		
Carbon	C^{IV}	12
Silicon	$\mathrm{Si}^{\mathrm{IV}}$	28·5
Platinum	$\mathrm{Pt}^{\mathrm{IV}}$	196·7
Tin in the stann*ic* state	$\mathrm{Sn}^{\mathrm{IV}}$	118

579. We shall in another chapter explain how atomic and molecular weights are determined.

Examination Questions.

435. Define 'element,' 'compound,' 'equivalent,' 'atom,' and 'molecule.'

436. How do you distinguish chemical affinity from the other molecular forces?

437. State the relation that generally exists between the sp. gr. of an elementary gas or vapour compared with hydrogen as unity, and its atomic weight.

438. How do you deduce the molecular weight of a gaseous or vaporisable compound from its sp. gr.? Give examples.

CHAPTER XI.

CHEMICAL NOMENCLATURE AND NOTATION—EXERCISES—BASIC SUBSTANCES—TABLE OF THE MORE COMMONLY OCCURRING BASES—BASIC SULPHIDES — EXERCISES — QUESTIONS — THE METHODS ADOPTED FOR PREPARING THE BASES—TABLE SHOWING THE SOLUBILITY OF THE BASES IN WATER AND THE ALKALIES—EXERCISES—QUESTIONS.

580. Nomenclature is the spoken language of chemistry, as notation is the symbolic written language of the science; and we have to explain the rules for naming and symbolising chemical compounds. We will commence with binary compounds, that is compounds which are composed of two elements only. The name of one of the elements in these compounds ends in *ide*, as ox*ide*, sulph*ide*, hydr*ide*, &c. If the compound consists of a metal and non-metal, it is the name of the non-metal that terminates in ide; and the name of the metal is now most frequently placed first, as potassium oxide, sodium sulphide, magnesium chloride, &c. When the two elements in combination are both non-metals, there is no exact rule as to which of the two names shall end in ide, but whichever name is made to end in ide it is placed last, as in the previous examples—thus, phosphorus chloride. Formerly the name of the non-metal was placed first. Examples: oxide of zinc, chloride of sodium, iodide of lead, &c.; and this method of naming binary compounds is not yet obsolete.

581. Many of the elements unite, as we are already aware, in more atomic proportions than one, forming several distinct chemical compounds: *prefixes* have therefore to be used to indicate the respective compounds. When one of the elements unites in six different proportions with the other element, say in the proportion of the numbers 1, 2, 3, 4, 5, and 6, the prefixes are *mono, di, tri, tetra, penta, hexa*; in place of these Greek prefixes the Latin prefixes *uni, bi, ter, quadro, quinque, sex* are often used; thus we can say potassium disulphide, or potassium bisulphide, both are equally correct; but the Greek prefixes are more generally employed at the present time. When the atomic relation between the two compounds is in the proportion of 2 to 3, the Latin prefix *sesqui*, meaning one and a half, is employed. The highest oxides, chlorides, sulphides, &c. are sometimes termed *peroxides, perchlorides*, &c. When the two elements are combined together in the proportion of one atom to one atom the prefix is frequently dispensed with, as, for instance, potassium chloride KCl.

[illegible] The student is already aware that when two distinct series of compounds are formed by the same elements, the name of one of the elements in the lower compound ends in *ous* and the name of the higher one in *ic*; and if one of the elements is a metal, it is the name of the metal that so terminates, thus ferrous chloride, ferric chloride. There is one rule to which we must draw the attention of the student, it is this, that some chemists make the name of the metal in a compound terminate in *ic* instead of in *um*, as potassic oxide, and [illegible] [illegible] [illegible].

Exercises on Nomenclature.

1. Name the following compounds: [illegible]

[illegible], [illegible], [illegible], [illegible], [illegible], HgCl, HgCl, [illegible] and [illegible].

[illegible] We will not proceed further at present with the nomenclature, but will commence with the notation. We already know that a small figure placed below the line multiplies only the element on the left-hand side of it, but this is not the case if the compound is enclosed in a bracket; if enclosed in a bracket the numeral indicates the number of molecules of a compound; thus, $3(HCl)$ denotes 3 molecules of the compound hydrogen chloride. The small figure may be placed either above or below the line, not upon it. A large figure placed upon the line denotes the number of molecules of a compound on the right of it; thus, $2Cr_2O_3$ denotes 2 molecules of chromium sesquioxide; the figure on the line multiplies all that follow, up to a *comma*, *full-stop*, or *plus sign*, not beyond these signs. When a compound is composed of two other compounds, it is usual to separate the molecules by a comma. Example: $2KCl,PtCl_4$. When a period is placed between compounds it indicates that they are in less intimate union than if they had been separated by a comma. 'For instance, in the formula for crystallised sulphate of magnesium and potassium, $MgSO_4,K_2SO_4.6H_2O$, the compound $MgSO_4$ is supposed to be more intimately united with K_2SO_4 than with the $6H_2O$, which may be expelled by heat.' The plus sign is sometimes employed in place of the full-stop. As a figure upon the line multiplies only up to a comma, full-stop, or plus sign, it is necessary, when we have to express more than one molecule of a substance, in the formula of which one or more of these signs have to be employed, to enclose the formula in a bracket, and place the number outside it, either before or after. Examples: $2(MgSO_4, K_2SO_4.6H_2O)$; $(ZnCl_2, 2NaCl)_2$.

Exercises on Notation.

440. Explain the following formulæ, the term formula being applied to a group of two or more symbols; the student is not expected to give the names of the compounds:
$Ba(NO_3)_2$. $ZnCl_2$, $2KCl$. $2(KI,HgI_2)$, $3(H_2O)$, $3(FeSo_4,7H_2O)$. $AlF_3,3NaF$. $Al(PO_4)$.

584. The student has already been informed that there are *four* different classes of compound substances, viz., *bases*, *acids*, *salts*, and *indifferent* bodies. We have now to make the student more fully acquainted with bases, acids, and salts, and we shall commence with the basic bodies.

Bases are bodies which unite with acids, and form with them the class of compounds termed salts. There are three classes of compounds to which the term base is applied. These are—

1st. Compounds composed of metals, or compound radicals playing the part of metals and oxygen. Example :—

Na_2O, [illegible] $Zn''O$, Fe'''_2O_3. The bases of this class are frequently called anhydrous bases.

2nd. Compounds composed of metals, or compound radicals playing the part of metals and the univalent radical hydroxyl (HO). Example: $NaHO$, NH_4HO, $Zn''H_2O_2$, $Fe'''_2H_6O_6$. The bases of this class are frequently called hydrated bases.

3rd. Compounds composed of hydrogen and nitrogen, and hydrogen and phosphorus. Example: NH_3.

585. The bases which are soluble in water *restore* the *blue* colour of litmus which has been *reddened* by an acid.

The following is a list of the more important bases belonging to the 1st and 2nd classes, and the student, before proceeding further, must commit to memory the names and formulæ of these bases. If the student does not follow my advice in committing to memory the names and symbols of the elements, and the names and formulæ of these bases and the other typical compounds which follow, he will encounter difficulties that he would not encounter if he had followed the advice given him.

586. The names placed in the brackets, in the following list, are the commercial and common names of the substances, as for example, caustic potash, caustic soda, quicklime, slaked lime, &c.

587. The student will not fail to notice that the hydrates of a few of the oxides are not given in the table: whether the oxides of these metals can form hydrates or not is uncertain; all that can be said is, they have not yet been obtained. The term *hydrated* was applied to the bases which contain hydrogen as well as oxygen when it was thought the whole of the hydrogen was united with a portion of the oxygen in the proportion to form water. The hydrated bases from this point of view are compounds of the oxides of the metals and water, and their formulæ, in accordance with this view, are thus expressed: Potassium hydrate, K_2O,H_2O, Barium hydrate, BaO,H_2O, Stannic hydrate, $SnO_2,2H_2O$, Ferric hydrate, $Fe_2O_3,3H_2O$, &c.; and this mode of representing this class of bases is still adopted by some chemical writers. Instead of the formulæ given n the table for this class of bases some chemical writers represent hem thus:—KHO, $Ba(OH)_2$,$Sn(OH)_4$,$Fe(OH)_6$, &c.; but, as e student will not fail to see, these formulæ, and those given

Sodium oxide (soda)	Na_2O	Sodium hydrate (hydrate of soda, caustic soda)	$NaHO$
Lithium oxide	Li_2O	Lithium hydrate (lithia)	$LiHO$
Ammonium oxide	$(NH_4)_2O$	Ammonium hydrate (ammonia)	$(NH_4)HO$
Argentic oxide	Ag_2O		
Barium oxide (baryta)	BaO	Barium hydrate (hydrate of baryta)	BaH_2O_2
Strontium oxide (strontia)	SrO	Strontium hydrate (hydrate of strontia)	SrH_2O_2
Calcium oxide (lime, quicklime)	CaO	Calcium hydrate (hydrate of lime, slaked lime)	CaH_2O_2
Magnesium oxide (magnesia)	MgO	Magnesium hydrate (hydrate of magnesia)	MgH_2O_2
Zinc oxide	ZnO	Zinc hydrate	ZnH_2O_2
Cadmium oxide	CdO	Cadmium hydrate	CdH_2O_2
Manganous oxide	MnO	Manganous hydrate	MnH_2O_2
Nickelous oxide	NiO	Nickelous hydrate	NiH_2O_2
Cobaltous oxide	CoO	Cobaltous hydrate	CoH_2O_2
Plumbic oxide (litharge)	PbO	Plumbic hydrate	PbH_2O_2
Cupric oxide	CuO	Cupric hydrate	CuH_2O_2
Ferrous oxide	FeO	Ferrous hydrate	FeH_2O_2
Ferric oxide, or iron sesquioxide	Fe_2O_3	Ferric hydrate	$Fe_2H_6O_6$
Chromic oxide, or chromium sesquioxide	Cr_2O_3	Chromic hydrate	$Cr_2H_6O_6$
Aluminium oxide (alumina)	Al_2O_3	Aluminium hydrate	$Al_2H_6O_6$
Mercurous oxide	Hg_2O		
Mercuric oxide	HgO		
Stannous oxide	SnO	Stannous hydrate	SnH_2O_2
Stannic oxide	SnO_2	Stannic hydrate	SnH_4O_4
Bismuthous oxide	Bi_2O_3		
Antimonious oxide	Sb_2O_3		
Platinic oxide	PtO_2	Platinic hydrate	PtH_4O_4
Auric oxide	Au_2O_3	Auric hydrate	AuH_3O_3

in the table, are really alike ; the only difference being that in the latter case the formula of the hydroxyl is placed in a bracket, and the number of molecules is indicated by a figure after the bracket. It is immaterial which plan the student follows.

588. The bases, both anhydrous and hydrated, are usually divided into the four following orders, viz., *alkalies*, *alkaline earths*, *earths*, and the *oxides* of the *heavy* metals.

589. The Arabians introduced the term *alkali* into chemistry; they discovered a substance possessing certain properties in the ashes of plants called *kali*, and to this substance they gave the name alkali by prefixing the Arabic definite article Al to the word kali. The alkalies given in the table are potash, soda, lithia, and ammonia ; there are some other alkalies besides these, but they are the oxides of some of the rare metals, and therefore they are not given in the list.

590. Potash and soda were, when first discovered, confounded together ; the ammonium base resembling them very much in properties was likewise termed an alkali. The ammonium alkali being *volatile* was called the *volatile* alkali ; and potash and soda *not* being *volatile* were called the *fixed* alkalies. The ammonium base was also termed the *animal* alkali, being derived from animal matter ; whilst soda was the *mineral* alkali, being derived from rock salk, or the ocean ; and potash received the name of the *vegetable* alkali, from its only known source at that time being the ashes of plants growing upon land.

591. The alkalies are very much more soluble in water than the next class of bases ; but they vary in solubility, lithia and its carbonate being more insoluble in water than the other three alkalies and their carbonates. The great solubility of the alkalies in water in comparison with the other bases is one of their distinctions ; they restore the *blue* colour of *reddened* litmus ; they turn the *yellow* of turmeric *brown* ; and they neutralise the strongest acids. The metals of the alkalies are termed alkaline metals, and NH_4, as it has the same functions as a metal, is called the compound metal ; it is as yet a hypothetical compound, for it has not been isolated.

592. Baryta, strontia, lime, and magnesia are termed alkaline earths : they are far less soluble in water than the alkalies; magnesia indeed being only very slightly soluble in water. They react faintly alkaline, that is to say they react on reddened litmus paper, and on the colour of turmeric like the alkalies, but more faintly; and they neutralise acids. They are distinguished from the alkalies by their carbonates being *insoluble* in water. Their metals are called the metals of the alkaline earths.

593. The only oxide we have given in the list belonging to the order of earths is alumina; the other oxides belonging to this order are the oxides of some of the rare metals; they do not, therefore, come within the scope of this work. The other earths, as well as alumina, are insoluble in water, and have therefore no action on reddened litmus or on the colour of turmeric.

594. The remaining basic oxides in the list are sometimes termed the oxides of the heavy metals; many of them are coloured, and are easily fusible or volatile; they may be said to be insoluble in water, although some of them exhibit a certain degree of solubility in that liquid, alkaline reaction, and metallic taste. Their affinity for acids is in some cases greater, in others less, than that of the earths. A few of them, as the auric and platinic oxides, have almost equal claims to be classed as acids, as they combine with strong bases.

595. The heaviest metals of this class of oxides are called *noble* metals, as mercury, silver, gold, platinum, and some of the rare metals. As regards the term heavy applied to metals, we may observe, it was universally believed at the commencement of the present century that every metal must possess a comparatively high sp. gr.—that is, must be considerably heavier than water. But on the discovery of potassium and sodium this view was found to be incorrect, as they are lighter than water. And lithium, which was isolated later, is the lightest known solid.

596. We requested the student to commit to memory the names and formulæ of the anhydrous and hydrated bases, not only for the reason that they are a most important class of compounds, and are therefore being constantly met with and alluded to in chemical writings; but also because they are the *types* of the composition of the *basic* sulphides, selenides, and hydrogen, on account of its weak basic character, has been combined with only two, hydrobromic and hydriodic, acids. There are two other hydrogen compounds which have a close affinity with the two we have just described; these are arsenuretted hydrogen (H_3As), and antimonuretted hydrogen (H_3Sb); these two hydrogen compounds have not been combined with acids, but when the hydrogen has been replaced by compound radicals they behave with acids like the similar compounds of nitrogen and phosphorus. Nitrine is the new name for ammonia, phosphine for phosphuretted hydrogen, arsine for arsenuretted hydrogen, and stibine for antimonuretted hydrogen.

Examination Questions.

445. What are the characteristics of a basic body?

446. Why is the formula of aluminium chloride notated Al_2Cl_6 and not $AlCl_3$?

447. Why have the bases that are insoluble, unlike those that are soluble, in water no action on paper stained with turmeric or with reddened litmus?[1]

601. By means of the following table the student will learn not only the bases that are soluble in water; but which of the bases, insoluble in that liquid, are *soluble* or *insoluble* in the fixed alkalies, and in the volatile one. This table is given as we have now to make the student acquainted with methods for obtaining the anhydrous and hydrated bases; and from the experiments and descriptions he will also become acquainted to some extent with the individual properties of the different bases.

602. The bases which are soluble in water, and not volatile, can be obtained in their anhydrous state in two ways:—1. By burning the metals in dry air or oxygen gas. 2. By decomposing those of their salts the acids of which are volatilised by heat. Nitrates and carbonates are the salts particularly employed for this purpose. K_2O and Na_2O are prepared by the first method; Li_2O, BaO, and SrO are prepared by igniting their nitrates; CaO and MgO are prepared on a large scale by igniting their carbonates. All my readers must have seen limestone, $CaCO_3$, converted into lime (quicklime) CaO by burning the $CaCo_3$ in a lime-kiln.

603. These bases can be obtained in the hydrated states by

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[1] *The rest of the chapter must be passed over for the present; directions will be given when the student is to study the remaining portion.*

Table showing the Solubility or Insolubility of the Basic Oxides and their Hydrates in Water, and the Alkalies.

Soluble in water.

NAMES	FORMULÆ	NAMES	FORMULÆ
Potassium hydrate (white) .	KHO	Barium hydrate (white) . .	BaH_2O_2
Sodium hydrate (white) .	NaHO	Strontium hydrate (white) .	SrH_2O_2
		Calcium hydrate (white) .	CaH_2O_2

The following are insoluble in water: lead, silver, and magnesium oxides are to a certain extent exceptions, as they dissolve slightly in water.

Soluble in ammonia and the fixed alkalies.

Zinc hydrate (white)			ZnH_2O_2

Insoluble in ammonia and the fixed alkalies.[1]

Ferric hydrate (reddish brown)	$Fe_2H_6O_6$	Mercurous oxide (black) forms no hydrate .	Hg_2O
Bismuthous hydrate (white) .	Bi_2O_3,H_2O	Mercuric oxide (yellow and red) forms no hydrate .	HgO

Insoluble in ammonia, soluble in the fixed alkalies.

Aluminium hydrate (white) .	$Al_2H_6O_6$	Chromic hydrate (bluish green). (This oxide is insoluble in boiling solutions of the fixed alkalies) .	$Cr_2H_6O_6$
Stannous hydrate (white) .	SnH_2O_2		
Stannic hydrate (white) .	SnH_4O_4		
Antimonious oxide (white) .	Sb_2O_3		
		Plumbic hydrate (white). (This hydrate is only very slightly soluble in the fixed alkalies)	PbH_2O_2

Soluble in ammonia, insoluble in the fixed alkalies. The presence of ammoniacal salts prevents some of them from being completely precipitated by the fixed alkalies.

Cobaltous hydrate (pale red) .	CoH_2O_2	Cupric hydrate (whitish green). (If the fixed alkalies are added to cold solutions of copper salts, the hydrate is precipitated; if added to boiling solutions, the anhydrous oxide is precipitated	CuH_2O_2
Nickelous hydrate (green) .	NiH_2O_2		
Argentic oxide (brown) forms no hydrate . . .	Ag_2O		
Cadmium hydrate (white) .	CdH_2O_2		

Insoluble in ammonia and the fixed alkalies, but in the presence of salts of ammonia the volatile alkali cannot precipitate them, and the fixed alkalies only do so in part.

Magnesium hydrate (white) .	MgH_2O_2	Ferrous hydrate is of a white colour, which, on exposure to the air, finally becomes red, owing to its being converted into ferric hydrate .	FeH_2O_2
Manganese hydrate (white), speedily becoming brown by absorbing oxygen from the air, and becoming converted into a higher oxide	MnO,HO		

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[1] Ammonia does not precipitate from mercurous and mercuric salts the oxides either in the anhydrous or hydrated state, but *mercurammonium salts*; the fixed alkalies likewise precipitate these mercurammonium salts, if salts of ammonia are present, but in the absence of ammonium salts they precipitate the oxides in the anhydrous state.

treating their anhydrides with water: we all know that calcium hydrate, slaked lime, is obtained in this way on a very large scale; and the union of the two compounds, we also know, is attended with a great evolution of heat. KHO and NaHO are prepared on the large scale by dissolving their carbonates in water, and adding to the solution slaked lime; the lime removes the CO_2 by combining with it, forming insoluble carbonate of lime; the KHO or NaHO, as the case may be, remains in solution. BaH_2O_2 is prepared on a large scale by heating $BaCO_3$ in earthenware retorts and passing superheated steam over it; the following decomposition takes place—

$$BaCO_3 + H_2O = BaH_2O_2 + CO_3.$$

MgH_2O_2 can be obtained by adding to a magnesium salt a solution of caustic potash or soda, as shown in the table.

604. The bases soluble, and most of those insoluble, in water, on being set free from their salts in aqueous solutions, combine with one or more molecules of water, the hydrated base being of course formed.

605. The term *caustic*, when applied to potash and soda, as caustic potash, &c., signifies the *hydrates* of these bases; but the term *caustic* is *only applied* to baryta, strontia, lime, and magnesia in their *anhydrous* state.

606. We will direct the student's attention to the fact that anhydrous potash, soda, baryta, and strontia cannot be obtained from their hydrates by heat, as water cannot be expelled from them by simple ignition; we shall have to call his attention to this fact again presently.

607. The bases insoluble in water can be obtained in their anhydrous state by several methods:—1. Many of the metals of these oxides unite with oxygen at high temperatures; when the metals are easily volatile they burn with a bright flame in the air or in oxygen. A few metals, such as gold and platinum, cannot be made to unite directly with oxygen even at the highest temperatures. The oxides of such metals can, however, be prepared by indirect means. Chlorine attacks all metals, and if the chlorides of these metals, which are not directly oxidisable, be decomposed by an alkali, their oxides are formed, thus:—

$$2AuCl + 2KOH = Au_2O + 2KCl + H_2O.$$

2. By decomposing those of their salts the acids of which are volatile, by heat.

3. They are not unfrequently prepared by expelling water from their hydrates by heat.

608. Most of these oxides, as shown in the table, form hydrates; and they are always obtained in their hydrated state by decomposing some *one* of their soluble salts when in a state of solution by an alkali or an alkaline earth; an alkali is preferred in the generality of cases to the latter. In the table their solubility or insolubility in the fixed alkalies and ammonia is shown. Now if a salt is decomposed by an alkali—volatile or fixed—in which the oxide set free from the salt is *soluble*, *none* of the liberated oxide will be *dissolved* if *just sufficient* alkali is added to combine with all the acid radical of the salt; if *less* of the alkali is added than is *sufficient* to combine with the *whole quantity* of the acid radical, then a *portion* of the salt will remain undecomposed; but if a *larger* quantity of the alkali is added than is necessary to *combine* with the acid radical, then some of the insoluble oxide will be *dissolved* in the *excess* quantity of the alkali, and if a *sufficient* excess is added, the *whole* of the insoluble base will be *dissolved* in the excess quantity.

609. Some of the hydrates possess a different colour to what the same oxides do when anhydrous.

610. We have now to consider the third class of bases, especially ammonia or nitrine; and with nitrine we shall connect ammonium, ammonia (NH_4HO), and hydroxalamine (NH_3HO).

611. NH_4 has not yet been obtained in the free state, but it exists as the basic radical in ammonium salts. It is a monad, and it is thought that if it could be obtained in the free state, it would possess metallic lustre, and the other characteristics of a metal, for it forms an amalgam—an unstable one—with mercury, which has a metallic lustre like all other amalgams. One method of preparing this amalgam is by pouring into a slightly warmed solution of NH_4Cl an amalgam of potassium or sodium; KCl or NaCl is formed and the ammonium amalgam, which rapidly decomposes into Hg, NH_3, and H. If I were teaching a class of students they might very naturally say to me that the formation of the ammonium amalgam was satisfactory, as far as it goes, to prove that the radical NH_4 exists in the ammonium salts; but why, as stated by chemists, this radical, if it exists in the ammonium salts, has not been obtained in the free state, like the simple metals to which it corresponds in functions, would be a very fair question for any student I was teaching directly to ask me. It is a still more pertinent question to ask by those who are only taught by me through the pages of this book, and I feel bound to answer it. In decomposing any ammonium salt by a metal, the NH_4, when set free, decomposes at the moment of liberation into NH_3 and H; possibly hereafter by conducting the decomposition at a lower temperature, or under greater pressure, or under both these conditions, than has yet been done, it may be obtained in the free state. The anhydrous oxide $(NH_4)_2O$, and the hydrated oxide (NH_4HO) have neither of them been obtained in their free state. When a solution of ammonia, which is considered to be a solution of NH_4HO, is exposed to the air, NH_3 escapes from it, even at common temperatures, and still more rapidly at higher temperatures; also, when an ammonium salt in solution is decomposed by another base, NH_3 is set free and not $(NH_4)_2O$, thus,

$$(NH_4)_2SO_4 + 2KHO = NH_3 + K_2SO_4 + 2H_2O.$$

As NH_3 escapes, like HCl, from an aqueous solution at common temperatures, an aqueous solution of ammonia was long considered to be, like a solution of hydrochloric acid gas in water, merely a mechanical solution of NH_3, no chemical union taking place between the NH_3 and H_2O; especially as the odour and other properties of the solution are the same as those possessed by the gas NH_3. But this view is no longer held; it is now considered that the ammonia solution is a solution of NH_4HO. Amongst other reasons for this view, all the four atoms of H in ammonium have been replaced by compound radicals having the same functions as H. Ammonia occurs in small quantities in the air, and it is present in rain water and in fertile soils; it is formed during the decay of animal and vegetable substances containing nitrogen; and some organic substances containing nitrogen when boiled in a solution of caustic potash are decomposed, the nitrogen being converted into NH_3. We have learned that it is evolved on decomposing ammonium salts by other bases; and that when it is passed into water, it is dissolved and converted into NH_4HO. *Hydroxylamine*, or *Oxyammonium*, NH_3O; this substance is a base, and combines, like NH_3, directly with acids forming crystallisable salts. It is formed by the direct union of H with NO; it is very volatile, and is easily decomposed, and can be obtained only in solution. Its salts are decomposed by KHO, with evolution of N and formation of NH_3.

612. *Phosphine. Arsine.*—Three compounds of P and H have been formed, but we shall only notice the one having the composition PH_3, and we only notice this on account of its basic character. It is a colourless gas, having an odour resembling that of garlic; it is very poisonous; it is combustible both in air and oxygen. It unites directly with HCl, HBr, and HI, forming compounds analogous to the haloid salts of ammonium, phosphonic bromide, PH_4,Br, for example. It can be prepared by several different methods; one of the easiest is to treat calcium phosphide with water :—

$$Ca_3P_2+6H_2O=2PH_3+3CaH_2O_2.$$

Arsine is a colourless gas, and is one of the most poisonous substances known. When exposed to a low red heat, without access of air, it is decomposed into H and As. It has not been combined with acids, but when the hydrogen is replaced by compound radicals, these compounds are basic. It is produced in all cases when nascent hydrogen and arsenic are brought together, especially in acid liquids; see exercise 717. This conversion of the As in its compounds into AsH_3 is employed for detecting minute traces of As; it is called Marsh's test.

613. On account of the basic hydrates being formed by the action of water on the anhydrous bases, it was long held by chemists in all countries that they were therefore compounds of water and the metallic oxides; but as far back as 1834, the late Mr. John J. Griffin, the well-known manufacturer of scientific apparatus, made the following statement with regard to these compounds in a book he published in that year. 'Every chemist but myself assumes the hydrates of metallic protoxides to consist of metallic protoxides combined with water. They make this assumption even in cases where the hydrates have been exposed, under common atmospheric pressure, to a *red heat*, without suffering decomposition, or giving off a particle of steam. But my opinion is that the hydrates of the metallic protoxides contain neither protoxides nor water, but are constituted of one atom each of metal, oxygen, and hydrogen. This is the *ultimate* constitution. What the *proximate* constitution is, I cannot tell. It may be $KH+O$, or $KO+H$, or $K+HO$; we have no means of determining *which*. I rest upon the ultimate constitution.' Some years after Mr. Griffin had published his views, the late eminent French chemist M. Gerhardt published similar views, and employed the same line of argument as Mr. Griffin had done.

614. We will now give a summary of the action which bases exert on salts.

615. When a salt is brought into contact with an additional quantity of the same base which it already contains, no reaction ensues, unless the acid of the salt can form a salt more basic than the original one. If potash be added to a solution of potassium sulphate, and the liquid evaporated, the original sulphate crystallises out unchanged. In other cases, combination ensues: thus, potash added to a solution of potassium disulphate produces the neutral sulphate.

616. If the base added to a saline solution is different from that existing in the salt, the original salt is frequently decomposed, and a new one formed; and the decomposition is determined by circumstances analogous to those which cause the reaction of acids on salts.

1st. *Generally speaking, a soluble salt, in a state of solution, is decomposed, when the reacting base can form an insoluble salt with the acid of the salt.* If baryta be added to a solution of potassium sulphate, barium sulphate is precipitated, and caustic potash remains in the liquid. Baryta also decomposes potassium carbonate in a dilute solution, and barium carbonate is precipitated.[1]

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[1] We have learnt in par. 603 that caustic potash and soda are manufactured from their carbonates by removing the CO_2 by means of slaked lime, the CO_2 forming with the lime an insoluble salt.

The state of concentration of the liquids exerts great influence over these decompositions; for, if barium carbonate is boiled with a concentrated solution of caustic potash, a considerable quantity of carbon dioxide is abstracted from it, and potassium carbonate is formed.

If the base of the original salt is insoluble it is precipitated, as well as the insoluble salt which has been formed by the union of the base which was added with the acid in the salt employed.

2nd. *A soluble salt containing an insoluble base is decomposed on adding to its solution a soluble base.* The soluble base decomposes the salt by taking away the *acid oxide* from the *insoluble base*; the latter—viz., the insoluble base—being set free will consequently precipitate, on account of its insolubility in the liquid. The bases, when set free in this manner, combine, with one or two exceptions, with water; this compound of the base and water is termed the hydrate of the base, whatever base it may be.

a. Some of the insoluble bases are *soluble* in solutions of the soluble bases, as shown in the table, and we here add a few examples:—Oxide of zinc is *soluble* in solutions of the fixed alkalies and ammonia; ferric oxide is *insoluble* in the fixed alkalies and ammonia; alumina is *insoluble* in ammonia, but *soluble* in the fixed alkalies; cupric oxide is *insoluble* in the fixed alkalies, but *soluble* in ammonia; magnesia is *insoluble* in ammonia and the fixed alkalies, but the volatile alkali cannot precipitate it from its solutions when salts of ammonia are present, and the fixed alkalies can only do so partly when these salts are present.

b. It has already been stated that if, to a solution of any of the soluble salts of the insoluble bases, a soluble base is added, the soluble base will take away the acid from the insoluble base; the latter (viz., the insoluble base) being set free, will consequently precipitate, on account of its insolubility in the liquid. If a *soluble salt* of an insoluble base is decomposed by a *soluble base*, which dissolves that insoluble base, the *insoluble base* will be *completely precipitated* if the soluble base is added only in sufficient quantity to take away the acid oxide from the insoluble base; but if more of the soluble base is added than is necessary to remove the acid oxide from the insoluble base, then in *the excess* of the soluble base the insoluble base dissolves; as an example, add to a solution of zinc sulphate ammonia, drop by drop, until a precipitate is formed; if the addition of ammonia is continued after the formation of the precipitate, the precipitate will finally dissolve in the free, that is, in the excess of ammonia.

3rd. *An insoluble metallic oxide sometimes decomposes a salt formed by a base equally insoluble.* Thus, oxide of silver decomposes nitrate of copper in solution, precipitating oxide of copper; the decomposition is determined, in this case, only by the preponderating affinity of oxide of silver for nitric acid.

4th. *When the base of a salt is volatile, it is generally expelled by a more fixed base,* particularly when assisted by heat.

Experimental Exercises on the Preparation of the Anhydrous and Hydrated Bases.

448. The experiments ought to be made by each of the students; and whenever it is not otherwise stated, they ought to conduct the operations in test-tubes. When want of room, or other causes, prevents the student from performing the experiments for themselves, the teacher ought to perform them. Dissolve the solid substances in water,[1] in the proportion of one part by weight of solid to ten parts by weight of water; and dilute the acids and ammonia[2] in the proportion of one of acid or ammonia to four of water, and use these dilute solutions always, unless expressly directed to use the concentrated. Caustic soda and caustic potash can be used indiscriminately; and they may be used as obtained from the operative chemist's.

449. Mix quicklime and solid ammonium chloride together, and introduce the powder into a flask and apply heat; nitrine will be evolved and calcium chloride formed. This experiment exemplifies the rule stated at par. 616—4th.

450. Dissolve one part of potassium carbonate in about ten parts of water. Boil the solution, and then add to it slaked lime (which has been made into a thin paste with water) until the clear solution no longer effervesces on the addition of hydrochloric acid in excess. The lime removes CO_2, forming with it calcium carbonate, which, on account of its insolubility, precipitates; K_2O is set free, and instantly combines with water, forming caustic potash, which dissolves in the water. This experiment may be made in a flask; and exemplifies the rule stated at par. 616—1st.

451. Add to a solution of sodium sulphate a solution of barium hydrate; barium sulphate will be formed, which, owing to its insolubility, will precipitate, whilst the Na_2O, on being set free, will combine with water and form the hydrate which will dissolve in the liquid. This experiment exemplifies the rule stated at par. 616—1st.

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[1] Distilled water ought to be employed; but if that cannot be obtained, rain water must be used.

[2] The word ammonia used here and in the exercises means the solution of NH_4HO, and not nitrine.

452. Add ammonia to a solution of ferric chloride; a precipitate ($Fe_2H_6O_6$) will be formed, which will remain undissolved on the further addition of the alkali. This experiment exemplifies the rule stated at par. 616—2nd.

453. Ignite bismuthous nitrate or carbonate in a porcelain crucible, the acid radical will be expelled, and Bi_2O_3 will be left.

454. Add ammonia to a solution of aluminium sulphate (common alum can be employed); a precipitate ($Al_2H_6O_6$) will be formed, which will not be dissolved on the further addition of ammonia (par. 616—2nd.)

455. To a solution of chromic chloride add a cold solution of caustic soda, until the precipitate which first appears is redissolved, then boil the solution until the hydrate once more precipitates.

456. Ignite some dry nickelous hydrate in a porcelain crucible, water will be expelled, and nickelous anhydride left. Observe the difference in colour between the anhydride and the hydrate.

457. To a solution of alum add caustic soda, until the precipitate which first forms is redissolved, then add hydrochloric acid until the solution manifests an acid reaction, and finally ammonia in excess.

458. Add caustic soda to a solution of ferrous sulphate. Boil another portion of the solution of the iron salt with a few drops of nitric acid until it becomes peroxidised, which will be indicated by the solution becoming yellow; when this is attained, add caustic soda in excess. Observe the difference in colour between the two precipitates.

459. Take three portions of a solution of cupric sulphate; to one add ammonia until the precipitate, which is first formed, redissolves; to the second portion add caustic soda in the cold. Boil the third, and add to it caustic soda.

460. Ignite cobaltous hydrate or cobaltous carbonate out of contact with the air; in either case CoO will be left; but if either of them is heated in contact with air, Co_3O_4 will be formed. These examples are given to show the student that not in all cases is the anhydride obtained if the hydrate or the carbonate is heated in contact with air (par. 616—2nd). To a solution of magnesium sulphate add ammonia, which will precipitate a portion of the magnesia as hydrate. To a solution of the same salt add ammonium chloride and then ammonia: in this case the ammonia will not produce a precipitate, as the oxides which are insoluble in ammonia, and yet not precipitated by it in the presence of ammoniacal salts, form with these salts double soluble salts, from which combinations the ammonia cannot precipitate them (par. 616– 2nd, *a*).

461. Mix together a solution of a ferric salt and a solution of aluminium sulphate (common alum can be employed), add caustic soda in excess to the mixed solution; $Fe_2H_6O_6$ will be precipitated; boil and filter. To the filtrate add hydrochloric acid in excess, and lastly ammonia, to precipitate the alumina; this illustrates one of the modes of separating the insoluble bases from each other, and detecting the different bases.

462. Ignite cobaltous nitrate in a porcelain crucible; Co_2O_3 and not CoO will be left. This experiment is given to show the student that not in all cases is the anhydrous protoxide of the metal left on the ignition of nitrates.

463. Dissolve some ammonium sulphate in water, place it in a stoppered retort, then pour into the retort some solution of caustic soda or potash; after the addition of the alkali and the retort being connected with the condenser and the receiver in its place, heat is to be applied by means of a gas-lamp or other heat source, and three-fourths of the liquid distilled over.

The distillate will be found to contain, from the smell, ammonia; and it will of course turn *red* litmus paper *blue*, and turn *turmeric* paper *brown*. This experiment exemplifies the rule stated at par. 616—4th.

Examination Questions.

464. Give as many methods as you can for obtaining cupric oxide in the anhydrous state.

465. If you add to a soluble salt, the base of which is insoluble in water, a base which forms, with the acid oxide of the salt, a salt insoluble in water, what will be the result?

466. If you are directed to add to a salt a base in excess, or not in excess; what is meant? Illustrate your answer by showing, by means of chemical equations, the chemical changes that would take place on adding, and not adding, ammonia in excess to a solution of aluminium sulphate.

467. If I burn one ton of limestone, how much quicklime ought I to obtain?

468. If I had to decompose 1 cwt. of NH_4Cl with CaH_2O_2, how much of the latter substance must I employ, and how much NH_3 would be set free?

469. I wish to prepare 10 lbs. of BaO by igniting $Ba(NO_3)_2$. How much of the nitrate must I employ?[1]

CHAPTER XII.

CHARACTERISTICS OF AN ACID—ACIDS EITHER MONOBASIC OR POLYBASIC—A LIST OF ACIDS AND ACID ANHYDRIDES—NOMENCLATURE OF THE ANHYDRIDES AND ACIDS—GRAPHIC NOTATION OF THESE SUBSTANCES—QUESTIONS—THE METHODS FOR PREPARING THE ACIDS AND ANHYDRIDES, WITH A DESCRIPTION OF THEIR PROPERTIES—EXERCISES—ACTION OF ACIDS UPON SALTS—EXERCISES—QUESTIONS.

617. SOME of the substances belonging to the class of bodies termed *acids* were known at a very early period; and the word acid, which is equivalent in ordinary language to sour, indicates one of the properties of the *soluble* acids, for they have a sour taste. Lavoisier, after he discovered oxygen, found by experiment that several acids contained this newly discovered substance; he concluded, in consequence, that oxygen was a constituent of all acids, and that it was the acidifying principle. According to this view an acid is an *oxidised* body having a sour taste, reddens blue organic colouring matters—litmus for example—and neutralises alkalies. But this definition was shown, some time afterwards, to be too limited, as a large class of bodies became known which possess all the properties of acids, and do not contain oxygen as a constituent. Moreover, there are some bodies which, from their *insolubility* in water, have no sour taste or action on vegetable colours, yet from their behaviour with bases are regarded as acids ; and some bodies admitted to be acids are even unable to *neutralise* completely the *alkaline* reaction of the alkalies.

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[1] *The student must now pass on to par.* 639, *Chap. XII., and complete the study of that chapter.*

618. In accordance with the most recent views, an acid may be defined as a compound containing one or more atoms of hydrogen, *which become displaced by a metal* when the latter is presented to the compound in the form of a basic hydrate ; the compound produced by the metallic substitution is termed *a salt.* The hydrogen capable of being so displaced may be conveniently termed *displaceable hydrogen.*

619. Examples of the substitution of a metal for the hydrogen in acids :

1st Example : Substitution of potassium for the hydrogen in hydrochloric acid :

$$\underset{\text{Hydrochloric acid.}}{HCl} + \underset{\text{Potassium hydrate.}}{KHO} = \underset{\text{Potassium chloride.}}{KCl} + \underset{\text{Water.}}{H_2O}$$

2nd Example : Substitution of potassium for the hydrogen in nitric acid :

$$\underset{\text{Nitric acid.}}{HNO_3} + KHO = \underset{\text{Potassium nitrate.}}{KNO_3} + H_2O$$

3rd Example : Substitution of one atom of potassium for one of the atoms of hydrogen in sulphuric acid :

$$\underset{\text{Sulphuric acid.}}{H_2SO_4} + KHO = \underset{\text{Hydric potassium sulphate.}}{KHSO_4} + H_2O$$

4th Example : Substitution of two atoms of potassium for the two atoms of hydrogen in sulphuric acid :

$$H_2SO_4 + 2KHO = K_2SO_4 + 2H_2O.$$

620. The student must not conclude that the hydrogen in acids can only be displaced by a metal when the latter is presented to the acid in the form of hydrate, but it can also take place when the metal is in the form of a basic anhydride, and it can take place in many cases when the metal is in the metallic state.

621. Examples of the substitution of a metal for the displaceable hydrogen, when the metal is in the form of a basic anhydride:

$$Na_2O + 2HCl = 2NaCl + H_2O$$

622. Examples of the substitution of a metal for the displaceable hydrogen in acids when the metal is in the metallic state:

$$Zn'' + H_2SO_4 = Zn''SO_4 + H_2$$

623. Acids containing one atom of displaceable hydrogen are said to be *monobasic*; those containing two atoms of displaceable hydrogen are said to be *dibasic*; those containing three such atoms are said to be *tribasic*, &c. Acids of greater basicity than unity are frequently termed *polybasic acids*.

624. When it became admitted that there were acids that contained no oxygen, they were divided into two groups, viz. oxygen acids and hydrogen acids: but according to the definition just given, all distinction is abolished: but as it will be convenient hereafter to distinguish the acids formerly termed hydrogen acids from those which were termed oxygen acids we will term the first set acids containing *non-oxygenated* radicals, and the second set we will term acids containing *oxygenated* radicals.

625. In the list of oxygen acids, a list of the corresponding *acid anhydrides* is given. The word anhydride signifies destitute of water, and is applied here to the oxygen compound which remains after the removal of the displaceable hydrogen in combination with sufficient oxygen to form water. Examples:

(1) $2HNO_3 - H_2O = N_2O_5$
(2) $H_2SO_4 - H_2O = SO_3$
(3) $2H_3PO_4 - 3H_2O = P_2O_5$.

626. The name of the oxygen acids is generally formed by adding *ic* as the terminal to the name, or to an abbreviation of it, of the element united with the oxygen. And when the same element forms two acids with oxygen, the one containing the least oxygen terminates in *ous*, and the other in *ic*.

627. We have given both the old and new names of the acids containing non-oxygenated radicals; but we have not given the new and more systematic names of the oxygen acids, as those given are almost universally employed by chemists at the present time. The names in italics are the oldest of the names given, and they are the common names; and one of them, sulphuretted hydrogen, is still very frequently used in chemical writings. The CN in the formula for hydrogen cyanide represents the composition of the compound radical cyanogen, which can be obtained in its isolated state, and is analogous in its chemical relations to the non-metallic elements, especially chlorine; thus we learn that some compound radicals resemble the non-metals in their chemical relations, whilst others, like ammonium, resemble the metals. CN is scarcely ever employed as the symbol of cyanogen, Cy being employed in place of it.

628. The names and symbols of the following acids, and their anhydrides, must now be committed to memory.

LIST OF THE MORE COMMONLY OCCURRING ACIDS.

Acids containing non-oxygenated radicals.

New Names.	*Old Names.*	
Hydrogen chloride.	Hydrochloric acid (*muriatic acid*)	HCl.
Hydrogen bromide.	Hydrobromic acid	HBr.
Hydrogen iodide.	Hydriodic acid	HI.
Hydrogen fluoride.	Hydrofluoric acid	HF.
Hydrogen cyanide.	Hydrocyanic acid (*prussic acid*)	$HCN = HCy$.
Hydrogen sulphide.	Hydrosulphuric acid (*sulphuretted hydrogen*)	H_2S.

Acids containing oxygenated acid radicals.

Nitric acid (*aqua fortis*)	HNO_3
Chloric acid	$HClO_3$
Sulphurous acid	H_2SO_3[1]
Sulphuric acid (*oil of vitriol*)	H_2SO_4
Carbonic acid	H_2CO_3[1]
Oxalic acid	$H_2C_2O_4$
Chromic acid	H_2CrO_4[1]
Arsenious acid	H_3ASO_3

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[1] Sulphurous, carbonic, chromic, and arsenious acids have not yet been obtained; they are supposed to exist in solution, but suffer decomposition when their solutions are evaporated. Sulphurous acid has, however, been recently

Arsenic acid	H_3ASO_4
Tribasic phosphoric acid or orthophosphoric acid	H_3PO_4
Boric acid or boracic acid	H_3BO_3
Silicic acid	H_4SiO_4

Acid anhydrides or acid radicals.

Nitric anhydride, nitrogen pentoxide	N_2O_5
Chloric anhydride (not yet obtained)	
Sulphurous anhydride, sulphur dioxide	SO_2
Sulphuric anhydride, sulphur trioxide	SO_3
Carbonic anhydride, carbon dioxide	CO_2
Oxalic anhydride (not yet obtained)	
Chromic anhydride, chromium trioxide	CrO_3
Arsenous anhydride, arsenic trioxide	As_2O_3[1]
Arsenic anhydride, arsenic pentoxide	As_2O_5
Phosphoric anhydride, phosphorous pentoxide	P_2O_5
Boric or boracic anhydride, boron trioxide	B_2O_3
Silicic anhydride (*silica*)	SiO_2

629. The anhydrides or, as they are also termed, *acid-forming oxides*, were formerly regarded as the acids, and the bodies we now term acids were considered to be compounds of water and these acids: carbonic anhydride is even yet best known under its old name carbonic acid. The oxygen acids are now frequently represented as compounds of the univalent radical *hydroxyl*, HO, and an oxygenated radical peculiar to each individual acid. Exs.: nitric acid $NO_2(HO)$, sulphurous acid $SO(HO)_2$, sulphuric acid $SO_2(HO)_2$, phosphoric acid $PO(HO)_3$ &c. When there are two or three different varieties of an acid, as in the case of phosphoric and silicic acids, the common one is termed the ortho-acid; ortho-phosphoric acid means the trihydric phosphate, ortho-silicic acid means the tetrahydric silicate. The different salts of the ortho-acids are also distinguished by the prefix ortho-, as ortho-phosphates, ortho-silicates.

630. Some of the acids, such as sulphurous, carbonic, and chromic acids, are so readily split up, as we have already stated, into their anhydrides and water that they cannot be isolated from their solutions, as they immediately decompose into these substances.[2] Other acids, on the contrary, cannot be decomposed

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obtained in combination with water in the form of cubical crystals on cooling a solution of it to 0° C.; the compound thus obtained has the formula $H_2SO_3 \cdot 14H_2O$.

[1] This compound is the *white arsenic* of the shops.

[2] Sulphurous acid in combination with water of crystallisation, we have pointed out in the note on page 283, has recently been obtained.

into their anhydrides and water; the anhydrides in these cases are obtained from their salts. Some anhydrides, such as chloric and oxalic, are so unstable that they have not been obtained.

631. In some cases there are more than two acids in a series; the sulphur oxygen acid series for example; when only two more were known belonging to the series besides those ending in *ous* and *ic*, the prefixes *hypo* and *hyper*, derived from two Greek words signifying *under* and *above*, were, and are to some extent still, employed; thus the acid $H_2S_2O_3$ was named hyposulphurous acid, because it contained a less quantity of oxygen than sulphurous acid for an equal quantity of sulphur; in like manner the acid $H_2S_2O_6$ was named hypo-sulphuric acid. These names are now being replaced by others; thus, the first is now named thiosulphuric acid, and the other is called dithionic acid; *thio* and *thion* being derived from a Greek word signifying sulphur. The salt, sodium thiosulphite, is largely employed in the arts, and its old name, hyposulphite of soda, is still its commercial name.

632. We have shown by examples in par. 625 the decomposition of the oxygen acids into their anhydrides. In the decomposition of the monobasic and bibasic acids there can only be *one stage* in the decomposition, as is there shown; in the case of the monobasic acids we see that *two* molecules of the acid are required in the decomposition; and in the case of bibasic acids *only one* molecule of the acid is required.

In the decomposition of tribasic acids there can be three stages:—1st, in which two atoms of acid are required, and one third of the hydrogen in the form of water is removed, a *tetrabasic acid remaining*. 2nd, in which one atom of the acid is required, and two-thirds of the hydrogen is removed in the form of water, a *monobasic acid remaining*. 3rd, in which two atoms of acid are required, all the hydrogen being removed in the form of water, the acid anhydride remaining. Examples:—

(1) $2H_3PO_4 - H_2O = H_4P_2O_7$

Trihydric phosphate. — Tetrahydric phosphate (Pyrophosphoric acid.)

(2) $H_3PO_4 - H_2O = HPO_3$

Monohydric phosphate. (Metaphosphoric [illegible])

(3) $$2H_3PO_4 - 3H_2O = P_2O_5$$
Phosphoric anhydride.

The student will see that there can be two stages in the decomposition of a tetrabasic acid—1st, in which *one-half* the hydrogen is removed in the form of water, a *bibasic acid* remaining; 2nd, in which the whole of the hydrogen is removed, the acid anhydride remaining. Examples :—

(1) $$H_4SiO_4 - H_2O = H_2SiO_3$$
Dihydric silicate. (Metasilicic acid.)

(2) $$H_4SiO_4 - 2H_2O = SiO_2$$
Silicic anhydride.

633. Some chemists have abandoned the term anhydride; in place of this name, they call this class of bodies 'acid-forming oxides.' When they are so termed, the second name given in the list of these bodies is adopted, as nitrogen pentoxide in place of nitric anhydride, &c.

634. All acid anhydrides are more or less quickly converted into the corresponding acids by the action of water or of hydrates. In the dry state, or when dissolved in ether or other liquids which do not change them into acids, they are without action on litmus and other vegetable colours.

635. The attention of the student was directed in the last chapter to the fact that sulphur combines with the metals, and that some of these sulphur compounds are *basic*. We have now to point out to him that some of these metallic sulphides, and also some non-metallic sulphides, possess the character of *acids*.

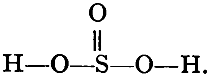
636. The following are all the sulphur compounds possessing acid characters that it is necessary to bring under the notice of the student. Thiocarbonic acid or sulphocarbonic acid, H_2CS_3, and the sulphur compound, CS_2, corresponding to carbon dioxide is called carbon disulphide, but it is better known in commerce by its old name bisulphide of carbon. Arsenic trisulphide, As_2S_3; arsenic pentasulphide, As_2S_5; antimony trisulphide, Sb_2S_3; antimony pentasulphide, Sb_2S_5; stannous sulphide, SnS; stannic sulphide, SnS_2; auric sulphide, Au_2S_3;

nic sulphide, PtS_2. We shall refer to these sulphides in the chapter.

›37. The anhydrides of the polybasic acids combine with ne, forming a peculiar class of acids. Examples:— hamic acid, NH_3SO_3; carbamic acid, NH_3CO_2; Oxamic , $NH_3C_2O_3$, &c.

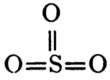
›38. We have now to direct the student's attention to the ing atomicity of some of the elements composing these pounds, and first with regard to sulphur. In the table of elements we represented it as a *dyad*, and it acts as a dyad H_2S; as a *tetrad* in sulphurous acid and its anhydride; *hexad* in sulphuric acid and its anhydride. These five hur compounds may, therefore, be graphically represented :—

) $H-S-H.$ (2) $O=S=O.$ (3)



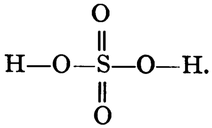
$$H-O-\overset{\overset{O}{\|}}{S}-O-H.$$

(4)



$$O=\overset{\overset{O}{\|}}{S}=O$$

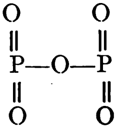
(5)



$$H-O-\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}-O-H.$$

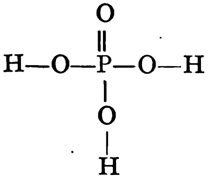
Nitrogen in nitric acid and its anhydride, phosphorus in phoric acid and its anhydride, and arsenic in arsenic acid its anhydride, are pentads. We show by two examples the e of formulating graphically the compounds of the pentad ıents when they contain the dyad element oxygen; we already given an example of the mode of formulating them they only contain monad elements :—

(1)



$$\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{P}}-O-\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{P}}$$

(2)



$$H-O-\overset{\overset{O}{\|}}{\underset{\underset{\underset{\underset{H}{|}}{O}}{|}}{P}}-O-H$$

Examination Questions.

70. Define the terms base and acid.

71. What is the valency of carbon in carbon monoxide (CO), and gen in nitrogen monoxide (N_2O)?

72. Explain what is meant by the terms monobasic, dibasic, and tri- acids, and give examples of each class.

473. Give the graphic formulæ of arsenous acid, and its anhydride.

474. Give the new names of hyposulphurous and hyposulphuric acids, and their formulæ.

475. Give the graphic formula of boracic acid.

476. What is the new name for hyposulphite of soda? and give its formula.[1]

639. We will now describe methods for obtaining the acids and acid-forming oxides, and also some of their properties.

640. *Hydrogen chloride. Hydrogen bromide. Hydrogen iodide. Hydrogen fluoride.*—The first three of these acids are gases, and they are soluble in water; the fourth is a liquid. Only HCl of these four acids is met with in nature, and it is met with in small quantities only, in certain volcanic gases. HCl is obtained by the action of H_2SO_4 on NaCl (Ex. 643) and HF by the action of H_2SO_4 on CaF_4. But H_2SO_4 cannot be usefully employed for the preparation of HBr and HI from their corresponding salts, for on their formation a portion is decomposed by some of the H_2SO_4:—

$$H_2SO_4+2HBr=Br_2+2H_2O+SO_2.$$

They can be obtained by heating one of their alkaline salts with H_3PO_4, and they can also be prepared by acting on their phosphides with water. They are very easily prepared in solution by suspending Br or I, as the case may be, and passing H_2S through the water, as both Br and I decompose H_2S by removing the H and combining with it. HCl in solution is manufactured on a very extensive scale, as the formation of Na_2SO_4 from NaCl is the first step in obtaining sodium carbonate and caustic soda from that salt: HCl is employed industrially, but it is more extensively used for the production of Cl (Ex. 673), the Cl being required in the manufacture of the well-known bleaching powder, chloride of lime. Cl decomposes both HI and HBr by removing the H and combining with it. HI in solution is rapidly decomposed by contact with air, the O removing the H.

641. *Hydrocyanic acid.*—This acid is a volatile liquid; it has a strong stupefying odour; it is present in laurel water, and in the water distilled from the kernels of many stone fruits, such as the plum, the peach, the apricot, &c. Although a valuable medicine in small doses, in larger doses it is one of the most powerful poisons known. It can be prepared from mercuric cyanide (HgCy) by adding to an aqueous solution of this salt H_2SO_4 diluted with water, and Fe in the state of filings ; on shaking the mixture and then distilling, HCy distils over :—

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[1] *The rest of the chapter must be passed over for the present: the student will be directed when to return to it.*

$$HgCy_2 + Fe + H_2SO_4 = 2HCy + FeSO_4 + Hg.$$

Chlorine easily decomposes it. Its compound radical CN is a colourless gas resembling the acid in odour. C and N do not unite together when heated alone, but they readily unite when K is present; thus, by passing pure N or air over an ignited mixture of carbon and K_2CO_3, KCy is formed.

642. *Hydrogen sulphide.*—This compound is a colourless gas having the odour of rotten eggs : if inhaled it has a poisonous action. It is evolved along with other gases from volcanoes, and is present in sulphur waters, as those of Harrogate. It is generally prepared by acting on metallic sulphides, as FeS, with H_2SO_4 ; there must be sufficient water present to dissolve the metallic sulphate formed, otherwise the H_2SO_4 will have little or no action on the sulphide. It burns in air or O with a bluish flame, if sufficient O is present H_2O and SO_2 are formed, but if there is an insufficient quantity only H_2O is formed, S being deposited. It is soluble in H_2O, from which it is expelled on boiling the liquid. It is in frequent use in the chemical laboratory for precipitating metals from their solutions as sulphides.

643. *Nitric acid.*—This is one of the acids that is manufactured on a large scale ; it was formerly obtained from potassium nitrate, saltpetre or nitre, as it is commercially called, by acting upon it with H_2SO_4 ; but since the discovery of sodium nitrate, or Chili saltpetre, as it is frequently called, in large deposits in the province of Tarapaca on the Peruvian coast, it has been substituted for the nitre, as it is cheaper and yields a larger proportion of HNO_3. The acid is a colourless liquid fuming strongly in the air, and absorbs moisture from the air with the greatest avidity. It is powerfully corrosive, cauterising the skin and staining it yellow. The acid of commerce contains about 30 per cent. of water, and boils at 120·5° C. It is

a powerful oxidising agent, and its action upon metals is of a somewhat complicated character, varying not only with different metals, but also, for the same metal, with the strength of the acid employed and the temperature at which the reaction takes place. Nitrates of the metals are formed, but at the same time another portion of the nitric acid is reduced to some lower oxide of nitrogen. All the metals in common use are oxidised by HNO_3 with the exception of Au and Pt, but they are not all dissolved ; Sn and Sb are left by the acid in the state of insoluble oxides. We are dwelling on this action of HNO_3 on the metals, because the action of this acid on the metals does not conform with the statement previously made that a metal can displace the hydrogen in the hydroxyl of the acid ; in all cases the NO_2 is partially or completely broken up and the metallic oxide formed is then converted into the nitrate by the action of some of the undecomposed HNO_3. Examples :—

(1) $Ag_2O+2HNO_3=2AgNO_3+H_2O$.
(2) $CuO+2HNO_3=Cu2NO_3+H_2O$.

We will now give three examples of the breaking up of the NO_2, or rather, as it should be said, of its deoxidation :—(1) When Ag is dissolved in the cold in an excess of HNO_3, nitrous acid (HNO_2) is formed :—

$$Ag_2+3HNO_3=2AgNO_3+HNO_2+H_2O.$$

When copper and other metals are acted upon by the acid of moderate concentration NO is evolved :—

$$3Cu+8HNO_3=3(Cu2NO_3)+2NO+4H_2O.$$

With a very concentrated acid NO_2 is evolved :—

$$Cu+4HNO_3=Cu2NO_3+2NO_2+2H_2O.$$

And when the decomposition occurs at a high temperature free N is usually disengaged in considerable quantities :—

$$5Cu+12HNO_3=5(Cu2NO_3)+N_2+6H_2O.$$

When Zn or Sn is acted upon by a very concentrated acid, not only is the NO_2 completely deoxidised, but NH_3 is one of the products formed :—

$$4Zn+9HNO_3=4(Zn2NO_3)+NH_3+3H_2O.$$

644. *Nitric anhydride.*—This is a white crystalline solid; it boils with decomposition at about 45° C. It can be prepared by acting on anhydrous HNO_3 with P_2O_5 :—

$$2HNO_3+P_2O_5=N_2O_5+2HPO_3.$$

It can also be prepared by passing Cl over dry silver nitrate. Two reactions occur in its preparation by this method; 1st, a volatile liquid called nitroxyl chloride, NO_2Cl, is formed; 2nd, this compound reacts on the excess of $AgNO_3$ with the formation of N_2O_5.

$$(1)\ AgNO_3+Cl_2=NO_2Cl+AgCl+O.$$
$$(2)\ NO_2Cl+AgNO_3=N_2O_5+AgCl.$$

645. *Aqua regia.*—We must notice here the mixture of HCl and HNO_3, to which the alchemists gave the name *aqua regia* from its power of dissolving gold, the *king* of metals. It is employed for dissolving Au and Pt, and also ores and other bodies. The solvent power of the mixture depends on the fact that, on heating the mixture, Cl is evolved, thus :—

$$HNO_3+3HCl=\underset{\substack{\text{Nitroxyl}\\\text{chloride.}}}{NOCl}+Cl_2+2H_2O.$$

646. *Chloric acid.*—This acid is a syrupy liquid of a yellowish colour; it is easily decomposed, yielding up its oxygen, hence it is a powerful oxidising agent. Organic bodies, as wool or paper, decompose it at once, also sulphur and phosphorus. It is prepared by acting on $Ba2ClO_3$ with dilute H_2SO_4; the reagents must be added in exact atomic proportions, as no excess of either must remain.

647. *Sulphurous anhydride.*—This compound is a colourless gas, which condenses to a mobile liquid when exposed to pressure or cold. Its suffocating odour is well known, as it is the compound always formed when sulphur is burnt in air or oxygen. It may be prepared not only by burning sulphur; but also by heating S and H_2SO_4 together; also by acting on Cu Hg, or Ag with concentrated H_2SO_4; it is made on a large scale, for the preparation of the sulphites, by heating carbon and H_2SO_4 together :—

(1) $S+2H_2SO_4=3SO_2+2H_2O$.
(2) $Cu+2H_2SO_4=CuSO_4+SO_2+2H_2O$.
(3) $C+2H_2SO_4=2SO_2+CO_2+2H_2O$.

648. *Sulphurous acid.*—This acid is formed by passing SO_2 into H_2O ; the solution has a strong acid reaction, and smells and tastes like the gas. Until recently the acid had not been isolated, owing to its so readily splitting-up into SO_2 and H_2O ; but it has recently been obtained on cooling its solution to 0°C. It crystallises out at that temperature, the formula of the crystals being $H_2SO_3+14H_2O$. Both SO_2 and the solution of H_2SO_3 exert a bleaching action on vegetable colours. The acid is employed in bleaching wool and silk.

649. *Sulphuric acid.*— This acid is manufactured on an enormous scale, on account of its many and important industrial applications ; the late Baron Liebig after describing some of the most important of them in his work 'Familiar Letters on Chemistry,' goes on to say : 'After these remarks you will perceive that it is no exaggeration to say, we may judge with great accuracy of the commercial prosperity of a country from the amount of sulphuric acid it consumes.' It is made on the large scale by conveying into large chambers made of lead dioxide or trioxide of nitrogen, steam, and SO_2 ; the following reaction takes place :—

$$SO_2+NO_2+H_2O=H_2SO_4+NO.$$

The NO is converted into NO_2 by uniting with some of the oxygen of the air in the chamber ; the NO_2 thus re-formed is again reduced to NO, owing to its yielding up one-half its oxygen to a fresh quantity of SO_2, and the NO is again converted into NO_2. This oxidation and deoxidation of these two oxides of nitrogen ought to go on indefinitely ; therefore, having once started a chamber with a sufficient amount of the oxide of nitrogen, no further amount ought to be required ; but in practice there is a considerable loss of the oxide, which has constantly to be replaced. Formerly the SO_2 was always made by burning sulphur, but now, except when the acid is required pure, it is obtained by roasting iron pyrites, FeS_2. The concentrated acid is a colourless oily liquid. The acid has a strong attraction for water, it withdraws it from the air, and chemists employ it for drying those gases on which it has no action, by passing them through it. So powerful is its attraction for water than when an organic substance, as wood, is dipped into it, it abstracts oxygen and hydrogen from the substance in the proportion to form water ; the organic substance becomes by this abstraction brown or black (charred), and the acid itself becomes brown. When it is diluted with water great heat is evolved. It is one of the strongest acids ; it expels nearly every other acid from its salts. Although the metals displace the hydrogen, a sulphate of the metal being formed, we have become aware that the concentrated acid when hot is decomposed by some of the metals. Au and Pt are unacted upon even by the boiling acid ; if its vapour is passed over red-hot platinum it is decomposed into SO_2, O, and H_2O. It has been proposed to employ this process for making oxygen on the large scale. Nascent hydrogen reduces the concentrated, but not the dilute, acid to H_2S. The acid boils at 620° F. and may be distilled without decomposition.

650. There is another sulphuric acid, viz. disulphuric or pyrosulphuric acid, $H_2S_2O_7$ (also called fuming sulphuric acid and Nordhausen sulphuric acid) ; it is obtained by heating a basic ferric sulphate in clay retorts.

651. *Sulphuric anhydride.*—This compound is a white crystalline solid ; it melts at 15° C. and boils at 46°. It is commonly prepared from Nordhausen sulphuric acid, by gently heating the acid in a retort ; the anhydride volatilises and condenses in the cooled receiver in white silky crystals of needle-like form. It is especially characterised by its strong attraction for water.

652. *Orthophosphoric acid.*—This acid is a solid, but as it deliquesces on exposure to the air, becoming a syrupy liquid, it is usually seen in that state. It is one of the strongest of the inorganic acids : it expels H_2SO_4 from its salts at high temperatures, as it is less volatile than that acid. It can be prepared by oxidising P with HNO_3, or by dropping PCl_5 in small quantities at a time into H_2O ; but it is prepared on the large scale by acting on bone-earth with sulphuric acid.

653. *Phosphoric anhydride.*—When P is burned in excess of dry air or oxygen this compound is formed ; it is a light snow-white powder; it readily attracts moisture from the air, becoming converted by its union with that liquid into phosphoric acid. Its attraction for water is so great that when thrown into that liquid it hisses like red-hot iron.

654. *Chromic anhydride and acid.*—The anhydride is prepared by adding to 1 volume of a cold saturated solution of potassium dichromate 1½ volumes of concentrated H_2SO_4; on the addition of the acid the mixture becomes strongly heated, on cooling the anhydride crystallises out in long red needles; the crystals are dried by placing them on a porous tile. It is a powerful oxidising agent, as it readily yields up half its oxygen, being converted into chromic oxide. It is readily soluble in water, and it is supposed to be converted by the solution into the acid; the acid, as previously stated, is unknown in the free state.

655. *Arsenious anhydride and acid.*—The anhydride is prepared on a large scale by roasting various minerals containing As. When first prepared it forms a colourless vitreous mass, which becomes after a time white and opaque; it volatilises at 380° F. It is employed in the manufacture of arsenical pigments, and in the manufacture of glass. The acid is probably contained in an aqueous solution of the anhydride, but it has not yet been obtained in the free state.

656. *Arsenic acid.*—This acid is obtained by treating As_2O_3 with HNO_3; on evaporating the solution hydrated crystals of the acid are deposited. It is, as we already know, a tribasic acid like H_3PO_4; like that acid it parts with, at different temperatures, hydrogen and oxygen in the proportion to form water, forming pyrarsenic acid, $H_4As_2O_7$, metarsenic acid, $HAsO_3$, and finally, at a low red heat, the anhydride, As_2O_5. The acid in solution gives up its oxygen to many substances even at ordinary temperatures. Nascent hydrogen reduces it, as well as arsenious acid, in acid liquids, to AsH_3. In consequence of the readiness with which it parts with oxygen, it is largely employed in the arts.

657. *Arsenic anhydride.*—As does not, like P, become converted into its highest oxide, As_2O_5, on being burnt in air or oxygen, because at the temperature of the combustion of arsenic As_2O_5 is decomposed into As_2O_3 and oxygen. This

anhydride is formed, as we have seen, on heating the acid to low redness ; it is a white amorphous substance.

658. *Boracic acid.*—The jets of steam which escape through the fissures in the ground in the volcanic districts of Tuscany contain traces of this acid ; the steam is made to pass through large basin-shaped vessels containing water, and the acid is left in the water. After the water has become charged with a considerable quantity of the acid it is evaporated to the crystallising point ; the acid crystallises out in colourless lustrous tablets. The aqueous solution of the acid possesses both a faint acid and alkaline reaction ; it reddens blue litmus paper, and yet it turns yellow turmeric paper brown. It has not yet been solved how the acid comes to be present in the jets of steam ; the most probable explanation appears to be, that it is produced by the action of the steam on boron nitride, as ammonium compounds are always present along with the acid in the steam. Boracic acid is a solid; it volatilises very slowly on intense ignition; it differs from most, if not all, other *solid* bodies, as it *volatilises* with the steam when its aqueous solution is boiled.

659. *Boracic anhydride.*—This compound is prepared by heating the acid to redness ; it is a colourless, transparent, vitreous solid.

660. *Silicic anhydride.*—This is the only oxide of silicon known ; it is very abundant in nature ; quartz or rock crystal is crystallised silica. 'The minerals *agate, chalcedony, opal, jasper, flint* consist of amorphous silica mixed to some extent with the crystalline modification. Some of these, *e.g.* opal, are of attractive colours and are used as jewels. Ordinary *sand* consists of particles of silica, which when united together by some cementing substance constitute the different varieties of *sandstone*.' It is a non-volatile solid, insoluble in water ; the only acid that acts upon it is HF.

661. *Silicic acid.*—This acid may be obtained in a gelatinous state by decomposing an alkaline silicate with HCl, or by passing CO_2 through the solution. It is difficult to obtain a silicic acid in the dry state of definite constitution, as, like other weak polybasic acids, it parts with water easily, finally becoming converted into the anhydride.

662. *Carbonic anhydride.*—This compound is always formed

679. *A normal salt is one in which the displaceable hydrogen of the acid is all exchanged for an equivalent amount of a metal or compound radical resembling in properties a metal.*

680. Having explained very fully the different varieties of salts that can be formed by the monobasic, bibasic, and tribasic acids, we need not enter upon the different classes formed by the quadrobasic acids, as the student will, or certainly ought to be able to, make them out for himself; but, in order to remove all reasonable difficulty, their general formulæ are given in the following table.

681. The nomenclature of salts is yet in a somewhat transitional state; and as, in the naming of the oxides, some chemists, as we pointed out, do not alter the name of the metal, whilst others add the termination *ic* to it, the same difference occurs in the naming of salts; thus, some chemists name, for example, KNO_3 and $ZnSO_4$, potassium nitrate and zinc sulphate, others name them potassic nitrate and zincic sulphate. We shall adopt, in the main, the first plan; but instead of writing the word hydrogen in full when that element is a constituent of a salt, we shall employ in its place, as is generally done, the word *hydric*. By the examples given in the following table, the student will see that we employ *no prefix* to express the quantity of the metal in a salt when the hydrogen of the acid is entirely replaced, not by different metals, but by *one* only; for instance, we do not designate Na_2SO_4 as disodium, but simply sodium sulphate: in like manner, we do not designate $Ca_3(PO_4)_2$ as tricalcium, but simply calcium phosphate; but it would not be out of order to call these salts disodium sulphate, and tricalcium phosphate. We have pointed out that when a metal has two different atomicities, the termination *ous* denotes the lower, and *ic* the higher, atomicity; this is followed out in the naming of the salts of such metals; thus, $Fe''SO_4$ is named ferr*ous* sulphate, and $Fe_3'''3SO_4$ ferr*ic* sulphate.

682. The general formulæ of the different classes of salts, which have been so far explained, are given in the following table, and the student, if he does not actually commit the formulæ to memory, must so study them that, if he were asked to write out a particular example of any one of them, he would

TABLE.—GENERAL FORMULÆ OF SALTS.

The letter M in the formulæ stands for the metal, and the dashes represent the atomicity of the metal.

GENERAL FORMULÆ OF SALTS.		EXAMPLES	
Nitrates	$M'NO_3$	Sodium nitrate	$NaNO_3$
	$M''(NO_3)_2$	Calcium nitrate	$Ca''(NO_3)_2$
	$M'''(NO_3)_3$	Bismuthous nitrate	$Bi'''(NO_3)_3$
Chlorates	$M'ClO_3$		
	$M''(ClO_3)_2$		
	$M'''(ClO_3)_3$		
Acid sulphites	$M'HSO_3$	Hydric potassium sulphite	$KHSO_3$
Sulphites	M'_2SO_3	Potassium sulphite	K_2SO_3
	$M''SO_3$	Calcium sulphite	$Ca''SO_3$
	$M'''_2(SO_3)_3$		
Acid sulphates	$M'HSO_4$	Hydric sodium sulphate	$NaHSO_4$
Sulphates	M'_2SO_4	Sodium sulphate Sodium ammonium sulphate	Na_2SO_4 $NaNH_4SO_4$
	$M''SO_4$	Barium sulphate	$Ba''SO_4$
	$M'''_2(SO_4)_3$	Aluminium sulphate	$Al'''_2(SO_4)_3$
Acid carbonates	$M'HCO_3$		
Carbonates	M'_2CO_3		
	$M''CO_3$		
	$M'''_2(CO_3)_3$		
Acid oxalates	MHC_2O_4		
Oxalates	$M'_2C_2O_4$		
	$M''C_2O_4$		
	$M'''_2(C_2O_4)_3$		
Acid chromates	$M'HCrO_4$		
Chromates	M'_2CrO_4		
	$M''CrO_4$		
	$M'''_2(CrO_4\ _3$		
Acid phosphates	$M'H_2PO_4$	Dihydric sodium phosphate	NaH_2PO_4
	M'_2HPO_4	Hydric disodium phosphate Hydric ammonium sodium phosphate	Na_2HPO_4 $NaNH_4HPO_4$
	$M''HPO_4$	Hydric magnesium phosphate	$Mg''HPO_4$
Phosphates	M'_3PO_4	Tripotassium phosphate	K'_3PO_4
	$M''M'PO_4$	Calcium potassic phosphate	$Ca''KPO_4$
	$M''_3(PO_4)_2$	Calcium phosphate	$Ca''_3(PO_4)_2$
	$M'''PO_4$	Bismuthous phosphate	$Bi'''PO_4$
Acid arseniates	$M'H_2AsO_4$		
	M'_2HAsO_4		
	$M''HAsO_4$		
Arseniates	M'_3AsO_4		
	$M''M'AsO_4$		
	$M''_3(AsO_4)_2$		
	$M'''AsO_4$		
Acid arsenites	$M'H_2AsO_3$		
	M'_2HAsO_3		
	$M''HAsO_3$		
Arsenites	M'_3AsO_3		
	$M''M'AsO_3$		
	$M''_3(AsO_3)_2$		
	$M'''AsO_3$		
Acid borates	$M'H_2BO_3$		
	M'_2HBO_3		
	$M''HBO_3$		
Borates	M'_3BO_3		
	$M''M'BO_3$		
	$M''_3(BO_3)_2$		
	$M'''BO_3$		
Acid silicates	$M'H_3SiO_4$		
	$M'_2H_2SiO_4$		
	$M''H_2SiO_4$		
	M'_3HSiO_4		
	$M''MHSiO_4$		
Silicates	M'_4SiO_4		
	$M''M'_2SiO_4$		
	M''_2SiO_4		
	$M'''MSiO_4$		
	$M'''_2M''(SiO_4)_3$		
	$'''_4(SiO_4\ _3$		

be able to do so at a moment's notice; and if he has followed the advice given him as regards committing perfectly to memory the names and symbols of the elements of the bases, and of the acids, and carefully studied the preceding explanations of the different classes of salts, he will be able most satisfactorily to accomplish any exercise of the kind; but a few exercises are given in order that he may put his knowledge into practice.

683. The formulæ of the following salts must be written out; and the molecular weight of two or three of each class of salts must be given in the way shown in the following example. The student will, of course, refer to the table of atomic weights for the weight of the different elements in the salts.

Example : Hydric disodium phosphate, Na_2HPO_4

$$\left\{\begin{array}{ll} Na_2 & = 46 \\ H & = 1 \\ P & = 31 \\ O_4 & = \underline{64} \end{array}\right.$$

Molecular weight of the salt 142

Exercises on the Notation and Molecular Weight of Salts.

497. Ammonium nitrate.
498. Barium nitrate.
499. Potassium nitrate.
500. Copper nitrate.
501. Lead nitrate.
502. Ferric nitrate.
503. Potassium chlorate.
504. Calcium chlorate.
505. Barium chlorate.
506. Calcium sulphite.
507. Hydric sodium sulphite.
508. Magnesium sulphite.
509. Strontium sulphite.
510. Ammonium sulphate.
511. Zinc sulphate.
512. Ferric sulphate.
513. Cobaltous sulphate.
514. Ferrous sulphate.
515. Copper sulphate.
516. Hydric potassium carbonate.
517. Barium carbonate.
518. Magnesium carbonate.
519. Sodium carbonate.
520. Hydric ammonium oxalate.
521. Nickelous oxalate.
522. Silver oxalate.
523. Cadmium oxalate.
524. Ferrous oxalate.
525. Silver chromate.
526. Lead chromate.
527. Sodium chromate.
528. Calcium chromate.
529. Potassium chromate.
530. Trisodium phosphate.
531. Magnesium ammonium phosphate.
532. Silver phosphate.
533. Lead phosphate.
534. Ferric phosphate.
535. Hydric cuprum arseniate.
536. Dihydric sodium arseniate.
537. Silver arseniate.
538. Magnesium ammonium arseniate.
539. Hydric cuprum arsenite (*Scheele's green*).
540. Silver arsenite.
541. Hydric calcium arsenite.

Exercises on the Nomenclature of Salts.

542. $CaSO_4$. 543. FeF_2. 544. $(NH_4)NO_3$. 545. $Pb(NO_3)_2$. 546. Fe_2Br_6. 547. $KClO_3$. 548. $MnCl_2$. 549. $BaSO_3$. 550.

MgH_2S_2. 551. K_2CrO_4. 552. $CaBr_2$. 553. $SrCO_3$. 554. $(NH_4)HS$. 555. Na_2HASO_4. 556. $AsCl_3$. 557. Na_3BO_3. 558. $CaHAsO_3$. 559. Fe_2SiO_4. 560. CuH_2SiO_4. 561. BaI_2. 562. Fe_2Cl_6.

684. Formerly the salts we have become acquainted with were divided into two classes: salts containing oxygen as a constituent formed one class, and were termed *oxysalts*; the other class, being composed of a metal and the characteristic non-metal in a hydrogen acid, were termed *haloid* salts, 'haloid' being derived from a Greek word signifying the sea. This term was employed to describe them, because common or sea salt, which is so universally known, is the type of this class of salts. The non-metals chlorine, bromine, iodine, and fluorine are frequently called *halogens*, because the salts they form with the metals resemble sea-salt in composition.

685. When this classification was accepted the oxysalts were regarded as composed of basic anhydrous oxides and the acid anhydrides, and these latter were considered to be the acids. This view was based on one of the methods for the formation of oxysalts, viz. if an acid oxide and a basic anhydride are brought together under suitable conditions they unite and form a salt; thus if the vapour of sulphuric oxide (SO_3) is passed over heated barium oxide (BaO) they unite and form the salt barium sulphate. Instead, therefore, of representing this salt by the formula $BaSO_4$, as we have done, they represented it thus, BaO,SO_3, and called it sulphate of baryta. This view of the constitution of the oxysalts subsequently gave place to what was termed the *binary theory* of salts. Both classes of salts are viewed by this theory as compounds of metals and acid radicals, the only difference being that in the oxygen salts the acid radical is always a compound, whereas at the time the theory was proposed the acid radicals in the haloid salts best known were elementary bodies.

686. There are a class of oxysalts which are termed *basic salts* or *subsalts*, because they contain in addition to the normal amount of the metal, one or more molecules of the basic hydrate of the metal: thus, dibasic lead nitrate, $PbNO_2.PbH_2O_2$; tribasic copper sulphate, $CuSO_4.2CuH_2O_2$. The bases of the monad metals, it may be stated, do not form basic salts. The haloid salts form a class of compounds by uniting with one or more molecules of the oxide of the same metal; for example, $PbCl_2.7PbO$. These compounds are called oxychlorides, oxyiodides, &c.

687. There are a class of salts which are the reverse of the basic salts in constitution, for they contain a larger proportion of the *acid oxide* than is sufficient to form a normal salt: these compounds are termed *anhydro* or *acid* salts. Exs.: sodium anhydrosulphate, $Na_2SO_4.SO_3$; potassium anhydrochromate, $K_2CrO_4CrO_3$, or potassium di- or bi-chromate, as it is generally called.

688. As there are, as we have shown, sulphur bases and sulphur acids, so there can necessarily be formed sulphur salts; some of these salts are known as potassium sulphocarbonate, or potassium thiocarbonate, K_2CS_3; potassium thioarseniate, K_3AsS_4, &c. As will be seen from these examples, sulphur salts resemble the oxygen salts in constitution, sulphur occupying the place of the oxygen. They have been much less studied than the oxysalts, and are, in a certain sense, of less importance than the oxysalts. The two rare elements, selenium and tellurium, which resemble sulphur in their properties, form corresponding salts.

689. Salts combine with each other, but by no means indiscriminately, forming what are termed double salts: these combinations can be most intelligently discussed in connection with the water of crystallisation.

690. Most of the coloured oxides form salts having various colours, thus :—

Ferrous salts are bluish-green.
Ferric salts are yellow or yellowish-brown.
Manganese salts are rose-coloured.
Chromium salts are dark green.
Nickel salts are green.
Cobalt salts are currant-red or blue.
Cupric salts are blue or green.
Gold salts are yellow.

No reason can be assigned for the particular colour of the salts of different metals, any more than we can for the different colours of the oxides of some of the metals: thus, for example, bismuthous oxide is *yellow*, and mercuric oxide is *red*. We should naturally expect that the salts of a coloured acid, like chromic acid, would be coloured; but even in this case, the neutral and basic salts of this acid with the same metal vary in colour: thus lead chromate ($PbCrO_4$) is *yellow*, whilst dibasic lead chromate ($PbCrO_4 \cdot PbO$) is *scarlet*. We have somewhat analogous examples in the difference in colour of the anhydrous and hydrated bases of the same metal; for instance, copper oxide is *black*, copper hydrate is *blue*.

691. We treated upon the solubility of substances in Chapter VI., but we could not there enter fully upon the *simple* or *mechanical solution* of substances; for, before we could do this, it was necessary for the student to have become fully acquainted with this third class of chemical compounds, viz. SALTS. We have therefore to re-enter upon the subject; and at the outset we have prominently to place before the learner the difference in the phenomena of solution, simple or mechanical, from those produced by chemical affinity. Chemical combination is uniformly attended with an evolution of heat, whilst solution is marked, with equal constancy, by cold. The phenomena of solution differ, also, 'essentially from those of chemical combination, inasmuch as in the former we have to do with gradual increase up to a given limit, termed the *point of saturation*, whereas in the latter we observe the occurrence of constant definite proportions, in which, and in no others, combination occurs. Solution obeys a law of continuity, chemical combination one of sudden change or discontinuity.'

692. In treating on the solution of salts, the only solvent of these substances we shall notice will be water; its solvent properties, we may remark, far exceed those of any other known liquid; and being the least costly liquid, it is always employed as the solvent except in particular cases. As it is such a universal solvent, the waters of all rivers, springs, and of the sea contain a varying quantity of certain salts, principally alkaline salts, and magnesium and calcium salts, which the solvent has dissolved out of the earth. Even rain water contains a small quantity of ammoniacal salt and some gases which the water has dissolved out of the air. In most chemical experiments therefore, distilled water has to be employed. 'The quantity of any solid, liquid, or gas which dissolves in water, or other solvent, must be ascertained *empirically* in every case, as we are unacquainted with any law regulating such solvent action, and we, therefore, cannot calculate the amount. And concerning the nature of solution, whether of solids, liquids, or gases, we know at present but little.'

693. Some salts are insoluble in water: the salts that are soluble in that liquid vary in their solubility; thus, one part of calcium sulphate is soluble in about 700 parts of water, while two parts of crystallised magnesium sulphate dissolve in 3 part of water. The solubility of most salts is increased by heating the liquid, but there are some salts that are more soluble in cold than in hot water; and there are some whose solubility increases with increase of temperature up to a certain point, but if the temperature of the liquid is raised beyond that point the solubility of the salt goes on decreasing as the temperature increases; thus sodium sulphate increases in solubility until the temperature of the liquid has reached 33° C.; if the temperature goes on increasing, the solubility goes on decreasing, until at the boiling-point there is retained in solution only about four-fifths of the quantity of salt which was dissolved at 33°. There is one salt at least, viz. sodium chloride, whose solubility is nearly equal at all temperatures. In the case of salts that are more soluble in hot than cold water, if a saturated solution of a salt of this kind be made at an elevated temperature and then allowed to cool, portions of the salt will, in most cases, go on depositing as the solution cools, until it reaches the ordinary temperature. The cold solution will be a saturated one, that is, will contain as much of the salt as cold water can dissolve. Hot saturated solutions of some salts, if not disturbed during cooling and no dust is allowed to fall in, will not deposit, during the cooling, any of the dissolved salt; but if the cold solution be shaken or be touched with a solid, or if even a particle of dust is allowed to fall into it, the crystallisation of the salt will at once commence, and in some cases the quantity of solid deposited is so great that the whole mass becomes solid. Such solutions are called *supersaturated* solutions.

694. Some salts, when they separate out of their aqueous solutions, have no water *chemically* united with them; such salts are termed *anhydrous* salts, and, as far as is known, these salts never form supersaturated solutions. Other salts, on the contrary, when they separate out of their aqueous solutions, are united with a certain definite number of molecules of water; and as these salts owe their crystalline character to the presence of this water—for when it is expelled, the salt ceases to be crystalline, and becomes an *amorphous* powder—it is hence termed *water of crystallisation*. The quantity of the water of crystallisation varies greatly in different salts; but each salt, as we have stated, always contains the same definite quantity under the same conditions. A salt may contain more than half its weight of water of crystallisation, and yet it will be a dry solid. In some salts this water is so feebly united with them, that it evaporates on the exposure of the crystals to the air, and on its evaporation the crystals lose their *lustre* and *shape*, and become *opaque*, and fall to *powder* or become *covered* with some of the amorphous white powder of the salt. Such crystals are said to *effloresce*; and the phenomena is called *efflorescence*. Although comparatively few salts effloresce, nevertheless the water of crystallisation can be expelled from the salt without the salt being *chemically* destroyed, for if we redissolve the salt, from which the water of crystallisation has been expelled, it will again combine with the necessary quantity of water for crystallisation and possess all its other distinguishing functions. The water of crystallisation is expelled, as a rule, on exposing the salt to a temperature of from 100° to 120° C.; but in some cases the salt requires to be exposed to a higher temperature to expel *all* the water of crystallisation: thus potash alum ($Al_2,3SO_4,K_2SO_4+24H_2O$), which contains twenty-four molecules of water of crystallisation, loses ten of the molecules at 100°, the second ten at 120°, but it retains the remaining four until the temperature reaches 200° C. When a salt, from which the water of crystallisation has been expelled, is redissolved in water, the solution is attended with a *development of heat*, owing to the salt again *chemically* uniting with the necessary quantity of water of crystallisation.

695. The reason of the solubility of sodium sulphate decreasing after the temperature of the solution rises above 33°, is due to the dissociation of *nine* of the *ten* molecules of the water of crystallisation; and a portion of this salt is

deposited as it forms because it is less soluble than the salt containing ten molecules of water; thus, below 33° the solution deposits crystals of the formula $Na_2SO_4.10H_2O$; above 33° C., the solution deposits crystals having the formula $Na_2SO_4.H_2O$. From this we learn that the temperature of the solution from which the crystals of some salts, which ordinarily crystallise with a large proportion of water, are deposited affects the quantity of the water of crystallisation with which the salt can combine; and the variation in the quantity of water alters the *crystalline form*; thus, 'in the case of sodium carbonate, whilst *monoclinic* crystals of the ten-molecule hydrate are deposited at the ordinary temperature, other crystals, having the composition $Na_2CO_3+8H_2O$, or again others represented by the formula $Na_2CO_3+5H_2O$, are deposited, when the crystallisation is allowed to take place at higher temperatures. On exposing a salt containing a large number of molecules of water of crystallisation to heat it undergoes what is termed *watery fusion.* A difference in the quantity of water of crystallisation *alters* the colour of a salt in some cases; thus anhydrous cobaltous chloride is *blue*, as is also the chloride having the following formula, $CoCl_2.2H_2O$; whilst $CoCl_2.6H_2O$ is *pink.*[1]

696. The late Professor Guthrie discovered that all salts, even those which are termed anhydrous, which are soluble in water form definite solid compounds with that liquid. The crystals of salts which are anhydrous at temperatures above 0° C., as NaCl and NH_4Cl, on crystallising out of an aqueous solution below this temperature contain water of crystallisation; thus a concentrated solution of common salt deposits at 7° crystals having the composition $NaCl2H_2O$; and at -23 crystals having the composition $NaCl10.5H_2O$. The hydrated salts formed below the freezing-point of water are termed *cryohydrates*; and each cryohydrate has a constant composition, and has a definite freezing and melting-point. Under certain conditions we learn, from Guthrie's discovery, that all salts are hydrated; therefore, when we speak of an anhydrous salt, we mean a salt that is anhydrous at temperatures above 0° C.

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[1] A dilute solution of this salt may be used as a sympathetic ink. When so employed it remains invisible when cold; but when the paper is placed before the fire it becomes blue and visible, and again fades in the cold, owing to its abstracting moisture from the air.

697. When two salts which do not chemically unite are dissolved in water, the quantity of each which is dissolved is, as a rule, less than if they were dissolved separately ; hence if to a concentrated solution of one salt another were added and the mixture shaken, some of the latter would dissolve, whilst a portion of the other salt would separate out of the solution. But some salts chemically unite with each other and form *double* salts; double salts are common between two different haloid salts ; for instance, between the haloid salts of the less oxidisable metals and the haloid salts of the alkaline metals. Exs. : Potassium platinochloride : $2\,KCl,PtCl_4$; potassio-mercuric iodide, $2KI,HgI_2$. The sulphates of the sesquioxides of the metals in many cases unite with the sulphates of the protoxides. Exs. : Potassium alum, $K_2SO_4,Al_2\,3SO_4.24H_2O$; ammonium alum $(NH_4)_2\,SO_4,\,Al_2\,3SO_4.24H_2O$; iron alum, $K_2So_4,\ Fe_23SO_4.24H_2O$, &c. Magnesium sulphate, zinc sulphate, and several sulphates of that class, contain seven molecules of water, their general formula being $MSO_4.\,7H_2O$, but one of these molecules is more strongly attached to the sulphate than are the other six, and it possesses quite different functions, as we shall show. If we expose crystals of one of these magnesium sulphates to a temperature of 100° C. six molecules are expelled, and with this loss the crystalline structure of the sulphate is destroyed ; but it requires a temperature of about 210° C. to expel the seventh molecule, and this seventh molecule is not water of crystallisation, for the crystalline structure was destroyed by the expulsion of the six molecules ; and, further, this seventh molecule can be replaced by some non-isomorphous anhydrous sulphate, as potassium sulphate. Ex. : If magnesium sulphate, or some other sulphate of that class, and K_2SO_4, be separately dissolved in equal molecular proportions, and the solutions mixed whilst hot, and allowed to crystallise, a double salt having the formula $MgSO_4,\,K_2SO_46H_2O$ is deposited, which has the same *crystalline* form as the magnesium sulphate, but differing in its constitution, in that the seventh atom of water in the magnesium sulphate has been displaced by potassium sulphate ; water that can thus be replaced by a salt has been termed *constitutional* or *saline* water.

698. The atomicity theory, which we have already described, the student will see does not explain why water chemically obtain them pure, they have to be redissolved in distilled water and recrystallised, sometimes three or four times. For the successful performance of the operation, considerable attention to minute detail is required, which can only be acquired by experience. When substances have the *same crystalline form*, crystallisation is not effectual for their separation, unless their difference in solubility is very great.

705. Many substances, both elementary and compound, crystallise in two or more forms belonging to different crystalline systems; such substances are said to be *dimorphous*, *trimorphous*, or, generally, *polymorphous*.

706. Some compounds which are analogous in chemical constitution, and are similar in crystalline form, are capable of replacing each other in the same crystals in *indefinite* proportions, when a mixed solution of these substances are submitted to the process of crystallisation; and there can also be obtained crystals consisting of alternate layers of different salts, if they have nearly the same degree of solubility in water. For example, if a crystal of potash alum ($K_2SO_4,Al_2,3SO_4.12H_2O$) be placed in a solution of ammonia alum ($(NH_4)\,2SO_4,Al_23SO_4.12H_2O$), the crystal augments in size owing to the depoposition on it of layers of the latter salt. If the crystalline mass be now transferred to a solution of iron alum ($K_2SO_4,Fe_23SO_4.12H_2O$) layers of this salt will be deposited on the crystal; if it then be transferred to a solution of chrome alum ($K_2SO_4,Cr_23SO_4.12H_2O$) layers of this salt will be deposited on it; and if it be now transferred to a solution of potash alum, layers of this salt will be deposited; the deposition of layers of the different alums could thus be continued again in succession. Substances which can thus replace each other in the same crystals in indefinite proportions are said to be *isomorphous*, a word which is derived from two Greek ones, one signifying *equal*, the other *form*.

707. From the preceding example we learn that the sulphates of aluminium, chromium, and iron in the ferric state, are isomorphous; also that the sulphates of potassium and ammonium are isomorphous. Isomorphous salts can never be separated by crystallisation, unless their difference in solubility is very great. Example: Copper sulphate and ferrous sulphate are isomorphous, and as they do not differ greatly in solubility they cannot be separated when present in the same solution by crystallisation; however frequently the substance is recrystallised the crystals always contain, throughout their whole mass, iron as well as copper. But if the iron is converted from the ferrous into the ferric state pure crystals of copper sulphate can then be obtained.

708. The same intermixture of isomorphous substances which takes place in artificial salts is found to occur in minerals, and affords a luminous explanation of the great variety, both in kind and proportion, of substances which may coexist in a mineral species, without its external character being thereby essentially affected.

709. Isomorphism frequently affords valuable aid in determining the atomic weights of the elements; for example, we only know *one* oxide of aluminium, which is composed of 18·3 parts by weight of the metal, and 16 of oxygen; is it a monoxide, a dioxide, a sesquioxide, &c., for upon the constitution of the oxide depends the atomic weight of the metal? As the aluminium oxide is isomorphous with sesquioxide of iron, it is also regarded as a sesquioxide; the atomic weight of the metal is therefore 27·4. Farther on we shall describe the method for determining the rational formula for compounds containing isomorphous constituents.

710. Nitrine unites, as has already been stated, with acids, and the salts formed are the same in character as those which are formed when ammonic hydrate unites with acids. Example:

$$NH_3 + HCl = NH_4Cl$$

Ammonium chloride.

$$2NH_3 + H_2SO_4 = (NH_4)_2SO_4$$

Ammonium sulphate.

711. Phosphine unites directly with hydriodic and hydrobromic acids when they are presented to it in the nascent state, forming compounds isomorphous with the corresponding salts in the nitrogen series:—

$$PH_3 + HI = PH_4I$$

Phosphonium iodide.

$$PH_3 + HBr = PH_4Br$$

Phosphonium bromide.

712. We noticed in the chapter on basic bodies that the hydrogen in ammonium can be replaced by metals ; there are consequently ammonium salts which contain a metal in place of some or all of the atoms of hydrogen in ammonium ; in some cases the metal replaces the hydrogen in a *single* ammonium molecule ; in other cases the metal, on its replacing its equivalent amount of hydrogen, *causes two* molecules of ammonium to *coalesce* into *one* molecule. Examples :—Dimercuric ammonium chloride, NHg_2Cl ; Mercuric diammonium chloride, $N_2H_6HgCl_2$.

Exercises on the Notation of Salts.

563. Potassium aluminium sulphate.
564. Cupric diammonium sulphate.
565. Sodium auric chloride.
566. Tribasic mercuric sulphate.
567. Platinic diammonium chloride.
568. Potassium ferric sulphate.
569. Mercuric ammonium chloride.

Examination Questions.

570. What is meant by the term 'isomorphism'?

571. What is meant by saying that a salt is efflorescent? and what by saying that a salt is deliquescent?

572. I dissolve a substance, say a salt, in water ; I add to the solution another liquid, say alcohol, which will mix with the solution ; after adding the liquid some of the solid dissolved in the water is precipitated ; what do you infer from that?

573. Explain what is meant by neutralisation, and what by saturation.

574. What is the difference between saline or constitutional water, and water of crystallisation?

575. I have some commercial potassium nitrate (saltpetre) containing potassium and sodium chlorides, which I wish to purify. In every 6,000 lbs. of the saltpetre these substances are present in the following proportions :

Potassium chloride	360 lbs.
Sodium chloride	840 ,,
Potassium nitrate	4,800 ,,
	6,000

1,200 lbs. of water, at a temperature of 212°F., is capable of dissolving the following quantities of these substances ; and the same quantity of water at 65° F. is capable of holding in solution the following quantities :

	At a temp. of 212° F.	At a temp. of 65° F.
Potassium chloride . . .	684 lbs.	396 lbs.
Sodium chloride . . .	324 ,,	318 ,,
Potassium nitrate . . .	4,800 ,,	348 ,,

I treat the 6,000 lbs. of saltpetre with 1,200 lbs. of boiling water. What takes place? I allow the *clear* boiling solution to cool down to 65° F. What occurs? and what finally remains in solution, and in what amounts?

576. I dissolve a substance in water; I add to the solution another liquid which will not mix with it, and being heavier than water sinks to the bottom. I shake the two fluids together, and then allow them to separate; when they have separated, I find the liquid has dissolved out of the water all the solid substance. What do you infer from that?

577. What is the difference between saline and crystalline water? Illustrate the difference in the formula of magnesium sulphate, and explain the part the former plays when this salt crystallises from the same solution with potassium sulphate.

578. What are cryohydrate salts? and what are anhydrous salts?

579. Does the atomicity theory explain why water chemically unites with salts? and why two salts unite together to form double salts?

580. On what does the action of freezing mixtures depend?[1]

713. We have now to bring under the notice of the student, (1) the action of metals on salts; (2) the action of salts upon each other; (3) the properties of each class of salts. For the student intelligently to understand much that will be stated, and which he will be called upon to perform and explain, the following table showing the solubility of different *normal* salts is required.

714. To ascertain the solubility of any salt by the table, find the name of the base in the upright column, and that of the acid in the line at the top; the number placed at the point where the two rows meet shows whether the salt formed by their combination is soluble or otherwise. Thus, if it is wished to find the solubility of a sulphide, hydrosulphuric acid is the acid under which it will be found; in like manner hydriodic acid is the acid for the iodides, hydrochloric acid the acid for chlorides, and hydrobromic acid the acid for bromides. The figures indicate the liquids which dissolve the substances. 'Thus, figure 1 means that it is soluble in water; 2, that it is insoluble in water, but soluble in hydrochloric or nitric acid; and 3, that it is insoluble in water and acids; 1–2 signifies a substance difficultly soluble in water, but soluble in hydrochloric or nitric acid; 1–3, a body difficultly soluble in water, and of which the solubility is not increased by the addition of the acids; and 2–3, a substance insoluble in water, and difficultly soluble in hydrochloric acid and in nitric acid. The deportment of a substance with hydrochloric acid differs materially from that which it exhibits with nitric acid in the following cases:—

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[1] *The student must pass over the rest of the chapter for the present; he will be directed when to return to it.*

Table showing the Solubility of Salts.

BASES ACIDS	Arsenious	Arsenic	Boracic	Carbonic	Chloric	Chromic	Hydrochloric	Hydriodic	Hydrosulphuric	Nitric	Oxalic	Orthophosphoric	Silicic	Sulphuric
Potash . .	1	1	1	1	1	1	1	1	1	1	1	1	1-2	1
Soda . . .	1	1	1	1	1	1	1	1	1	1	1	1	1-2	1
Ammonia . .	1	1	1	1	1	1	1	1	1	1	1	1		1
Magnesia . .	2	2	2	2	1	1	1	1		1	2	2	2-3	1
Lime . . .	2	2	2	2	1	1	1	1	1-2	1	2	2	2-3	1-3
Baryta. . .	2	2	2	2	1	2	1	1	1	1	2	2	2-3	3
Strontia . .	2	2	2	2	1	2	1	1	1	1	2	2	2-3	3
Aluminium oxide .	2	2			1	2	1			1	2	2	2-3	1
Chromic oxide .		1	2	2	1	2	1			1	1	2	2-3	1
Zinc oxide . .		1-2	2	2	1		1	1	2	1	2	2	2-3	1
Manganous oxide .		2	2	2	1	1	1	1	2	1	2	2	2-3	1
Ferrous oxide .	2	2	2	2	1		1	1	2	1	1-2	2	2-3	1
Ferric oxide. .		2	2		1	1	1	1	2	1	1-2	2	2-3	1
Nickelous oxide .	2	2	2	2	1	1	1		2	1	2	2	2-3	1
Cobaltous oxide .	2	2	2	2	1	2	1		2	1	2	2	2-3	1
Antimonious oxide	2	2				2	1-2		2		1-2			2
Mercurous oxide .	2	2	1	2	1-2	2	3	2	2	1-2	2	2		1-2
Mercuric oxide .	2	2		2	1	1-2	1	2	2	1	2	2		1
Plumbic oxide .	2	2	2	2	1	2-3	1-3	2	2	1	2	2	3	3
Cupric oxide .	2	2	2	2	1	2	1	2	2	1	2	2	3	1
Argentic oxide .	2	2	2	2	1	2	3	3	2	1	2	2		1-3
											2	2		1

1. The sulphides of cobalt and nickel are difficultly soluble in hydrochloric acid, but are readily decomposed by nitric acid.

2. Stannous and stannic sulphides are decomposed and dissolved by hydrochloric acid; by nitric acid they are each converted into stannic oxide, which is insoluble in an excess of the acid.

3. Sulphide of copper is difficultly decomposed by hydrochloric acid, but with facility by nitric acid.

4. Sulphide of silver is only soluble in nitric acid.

5. Platinic sulphide is not affected by hydrochloric acid; boiling nitric acid converts it into a soluble sulphate of the platinic oxides.

715. We have already learned that some metals can precipitate others from their salts, when the latter are in a state of solution, some of the *precipitating* metal combining with the acid radical which was in combination with the *precipitated* metal. Example:—

$$Cu + 2AgNO_3 = Cu(NO_3)_2 + Ag_2.$$

If a piece of iron be introduced into a solution of cupric sulphate the iron instantly becomes covered with a layer of metallic copper, precipitated by that portion of the iron which enters the solution.

716. The metal in an oxysalt is displaced by a more oxidisable metal, if the salt be in a state of solution. The action of metals on this class of salts, if they are in a state of solution, may therefore be foreseen; but it is difficult to generalise upon this action of metals upon metallic oxides *under all conditions*; for the relative affinity of the metals for oxygen varies greatly with temperature. Thus, potassium decomposes oxide of iron at a red heat, whilst at a higher temperature, as a strong white heat, iron, on the contrary, decomposes potassium oxide.

717. It may be further stated that in this respect the chlorides behave like the oxygen salts; iron, for example, displaces copper from cupric chloride as from the sulphate.

718. The following table indicates the order in which the metals precipitate other metals from their salts.

Salts in a state of solution reduced by certain metals.

Salts of tin		reduced by iron and zinc.
,, antimony		
,, bismuth		
,, lead		
,, copper	reduced by iron, zinc, and all those which precede copper	
,, mercury		
,, silver	reduced by iron, zinc, manganese, cobalt, and all those which precede silver	
,, platinum		
,, gold		

Experiments on the Action of Metals upon Salts.

581. To a solution of stannic chloride add a strip of iron ; the tin salt will be reduced by the iron.

582. To a solution of cupric sulphate add a piece of tin ; the copper salt will be reduced by the tin.

583. To a solution of mercuric nitrate add a piece of copper ; the mercury salt will be reduced by the copper.

719. *Action of salts upon each other.*—When two salts are mixed together, several phenomena may ensue. We shall distinguish the case in which a mixture of two salts in the *dry* state is heated, and that in which they are brought into contact in solution, or in the *humid* way, as it is termed.

720. *Mutual action of salts in the dry state.*—When two salts containing the same acid, but different bases, are heated together, the two salts frequently combine in definite proportions, producing double salts, which crystallise on cooling. In this manner a great number of double silicates may be produced, which, from their beautiful crystallisation, present the characters of definite compounds. In the same manner we may obtain, in the dry way, double chlorides and several other double salts ; but, so formed, the combination is often destroyed upon dissolving the compound in water, the two original salts crystallising separately.

721. *When two salts of different acids and bases are heated together, and when, by the mutual interchange of acids and bases, a new salt, more volatile than the first two, can be formed, its formation is generally determined by this circumstance.* Example :—Mix some calcium carbonate and ammonium chloride together in the dry state, and then expose the mixture to a temperature higher than 212° F., ammonium carbonate being a more volatile salt than the chloride, the two salts react upon each other, the following changes taking place.

$$CaCO_3+2NH_4Cl=(NH_4)_2CO_3+CaCl_2.$$

722. *Mutual action of Salts in the humid way.*—The solutions of two salts which have the same acid or base, are generally without action upon each other. Sometimes, however, the two salts unite and form a double salt. Aluminium sulphate, for instance, combines with potassium sulphate, forming a double salt, known by the name of *alum*. Potassium chloride combines with platinum chloride and produces a double chloride of platinum and potassium.

723. *Two soluble salts in a state of solution mutually decompose each other, when they can, by the interchange of the acids, form an insoluble salt, or a salt less soluble than either of the two original salts.*

724. It frequently happens that the chemical changes which take place in the dry way are the reverse of those which take place in the humid way. Example :—We have learned that calcium carbonate and ammonium chloride in the dry state decompose each other, when heat is applied ; but if we bring together in solution ammonium carbonate and calcium chloride, owing to the insolubility of calcium carbonate, double decomposition takes place, calcium carbonate and ammonium chloride being produced ; the former salt, being *insoluble*, is *precipitated*, whilst the latter, being *soluble*, remains in *solution*.

725. *An insoluble salt may sometimes be decomposed by boiling it for a long time with a soluble salt.* This occurs whenever the base of the original insoluble salt can form an insoluble salt with the acid of the reacting soluble salt.

To render the decomposition complete of the original insoluble salt, a large excess of the soluble salt must in the generality of cases be employed.

726. *We have learned that two soluble salts in a state of solution decompose each other, when they can by the interchange of the acid radicals form an insoluble salt, or a salt less soluble than either of the two original salts.* We can, therefore, by the aid of the table, not only ascertain whether any particular salt is soluble or insoluble, but also, if we add two soluble salts together in a state of solution whether any precipitation would take place, and if so what chemical changes must ensue. Example :—On adding an aqueous solution of ammonium carbonate and one of calcium chloride together, we should find, by

examining the table, that calcium carbonate was insoluble in water; these two salts would, therefore be mutually decomposed and two new ones formed, whenever aqueous solutions of the two salts were brought together.

727. We are already aware that if we bring a base, whether in the anhydrous or the hydrated state, in contact with an acid, they will form a salt, and by the aid of the table we can ascertain whether the salt would be soluble or insoluble; further, *if we add to a salt in solution an acid which forms with the metal in the salt an insoluble salt, the insoluble salt will be formed and precipitated, whilst the acid of the original salt will be set free.*

Experimental Exercises on the Formation and Decomposition of Salts.

Required to know whether any precipitation will occur when the following substances are added together; and if so, what chemical changes must ensue.[1]

583*a*. If an aqueous solution of cupric sulphate were added to one of lead nitrate?

584. If an aqueous solution of magnesium sulphate were added to one of barium nitrate?

585. If hydrochloric acid were added to an aqueous solution of mercurous nitrate, the solution of the mercury salt containing a small quantity of free nitric acid?

586. If an aqueous solution of potassium nitrate were added to one of calcium chloride?

587. If an aqueous solution of ammonium oxalate were added to one of calcium chloride?

588. If an aqueous solution of sodium chloride were added to one of lead nitrate?

589. If an aqueous solution of sodium phosphate were added to a hydrochloric acid solution of calcium chloride?

590. If ammonium sulphide were added to an aqueous solution of cupric sulphate?

591. If an aqueous solution of potassium chromate were added to one of barium chloride?

592. If an aqueous solution of magnesium sulphate were added to one of potassium nitrate?

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[1] The student must state whether any precipitation will occur, and if so, what chemical changes will ensue, before the experiment is made. He must also be made to generalise them; for instance, if he states that sulphate of lead will be precipitated when an aqueous solution of sulphate of copper be added to one of nitrate of lead, he must then say whether an aqueous solution of any soluble sulphate, on being added to a solution of nitrate of lead, would produce a precipitate of sulphate of lead; and then he must say whether sulphate of lead would be formed on adding a solution of any soluble sulphate to a solution of any soluble salt of lead; he must be made to generalise every question which admits of it in this way.

593. If an aqueous solution of ammonium carbonate were added to one of calcium chloride?

594. If an aqueous solution of sodium carbonate were added to one of ferrous sulphate?

595. If an aqueous solution of ammonium chloride were added to one of zinc sulphate?

596. If an aqueous solution of potassium chromate were added to a itric acid solution of barium nitrate.

728. We shall now briefly describe some of the chief general properties of the salts given in the table of solubilities; and a *test* for the acid oxide or halogen of the salt.[1]

729. *Chlorides.*—The chlorides are solid bodies with the exception of $SnCl_4$; some of them melt at comparatively low temperatures; and HgCl, $HgCl_2$ and $SbCl_3$ are also easily volatilised. HgCl and AgCl are the only chlorides which are completely insoluble in water, as will be seen by the table. A chloride, treated with sulphuric acid, disengages hydrochloric acid. When a chloride is heated with manganese dioxide and sulphuric acid, chlorine is evolved, which may be recognised by its odour and greenish-yellow colour. Chlorides dissolved in water give with silver nitrate a white precipitate, which assumes a violet tinge by exposure to light; the silver chloride is insoluble in acids but soluble in ammonia. Employ sodium chloride for these experiments.

730. *Bromides.*—The metallic bromides resemble the corresponding chlorides in properties. A bromide treated with sulphuric acid disengages hydrobromic acid, and also vapours of bromine, which impart a brown colour to the acid gas. If the bromide be heated with a mixture of sulphuric acid and manganese dioxide bromine only is disengaged Silver nitrate produces in solutions of the bromides a yellowish-white precipitate of silver bromide, which is insoluble in acids, but sparingly soluble in ammonia. Employ sodium or potassium bromide for these experiments.

731. *Iodides.*—The iodides in their general properties resemble very much the corresponding chlorides and bromides. The iodides treated with sulphuric acid instantly produce a considerable deposit of iodine; and if the mixture be heated, intense violet-coloured vapours are disengaged. The reaction is due to the decomposition of sulphuric acid by hydriodic acid, water and SO_2 being formed, and iodine set free. If the iodide be mixed with manganese dioxide, red lead (Pb_3O_4), or with dioxide of lead, upon heating the mixture with sulphuric acid, iodine is evolved, but no SO_2. Iodine, in a free state, forms with starch, even in highly dilute solutions, a purple precipitate of iodide of starch. If the iodine is in a state of combination with hydrogen or any metal, it is necessary to liberate it before applying the starch test. The iodine may be liberated from hydrogen or the metals by nitric acid, or chlorine ; if chlorine is added, an excess of it must not be employed, as it forms a colourless compound with iodine. Employ potassium iodide for these experiments.

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[1] The experiments indicated under each class of salts ought to be performed.

732. *Sulphides.*—Sulphur, like oxygen, frequently forms several compounds with the same metal, so that we may have monosulphides, disulphides, trisulphides, &c. A monosulphide, heated with dilute sulphuric acid, or with hydrochloric acid, disengages sulphuretted hydrogen, which is easily recognised by its odour, and no sulphur is deposited. A polysulphide also disengages sulphuretted hydrogen on being heated with dilute sulphuric acid, or hydrochloric acid ; but, in addition to the evolution of the sulphuretted hydrogen, a deposit of sulphur is formed. The evolution of sulphuretted hydrogen on the addition of hydrochloric acid, is at once characteristic of a sulphide ; and when the quantity of sulphuretted hydrogen is so minute that the smell fails to afford a sufficient proof, it may be detected by holding a piece of paper, moistened with a solution of any soluble salt of lead, over the mouth of the test-tube, as a brown or black coating of sulphide of lead will be formed upon the paper. When the sulphides of the heavy metals are heated in contact with air, SO_2 is evolved, the metal being left, in some cases, uncombined, and in others as an oxide. The sulphides of the alkalies and alkaline earths are converted by this process into sulphates. Oxidising agents convert the sulphides into sulphates. Employ some soluble sulphide, as sulphide of potassium or ammonium, for these experiments.

733. *Sulphates.*—Most of the sulphates, with the exception of the sulphates of barium, strontium, and lead, are soluble in

water, as will be seen by the table. The sulphates of the alkalies, alkaline earths, and sulphate of lead, are the only salts of this acid which are not decomposed on simple ignition; the other sulphates are decomposed, and generally yield a gaseous mixture of SO_2 and oxygen. Some sulphates, however, are decomposed at so low a temperature that the acid oxide is disengaged. All the sulphates are decomposed by carbon assisted by heat; the products vary with the nature of the base and the temperature. Sulphuric acid and the soluble sulphates give, in a state of solution, with a solution of a barium salt a white precipitate of barium sulphate, which, from its insolubility in acids, is at once distinguished from all other barium salts; any soluble barium salt in a state of solution is therefore the best and most delicate test for sulphuric acid. Employ a solution of any soluble sulphate for this experiment.

734. *Sulphites.*—The sulphites of the alkalies and alkaline earths are changed into sulphates and sulphides on being heated in a close vessel :—

$$4K_2SO_3 = 3K_2SO_4 + K_2S.$$

The other metallic sulphites disengage SO_2 and an oxide of the metal remains as a residue. Heated with carbon, they give products similar to those of the sulphates. Concentrated boiling nitric acid changes the sulphites into sulphates. Chlorine produces the same change in the sulphites in solution. The soluble sulphites also absorb oxygen from the air, and are changed into sulphates. The sulphites may easily be recognised by the characteristic odour of SO_2 which they emit when treated by hydrochloric acid, or by dilute sulphuric acid, which evolution is not accompanied by a separation of sulphur: employ sodium or potassium sulphite for this experiment.

735. *Nitrates.*—All the neutral nitrates, as will be seen by the table, are soluble in water; a few of the basic salts are insoluble in that liquid. Heat decomposes this class of salts, affording products which are rich in oxygen, and which powerfully assist combustion. In consequence of this property, the nitrates deflagrate on hot coals, and often detonate when heated with powdered charcoal. The alkaline nitrates, when subjected to a gradually increasing temperature, disengage at

first pure oxygen, and are changed into nitrites (MNO_2). Heated still further, they are entirely decomposed, evolving nitrogen and oxygen. The other nitrates disengage oxygen and NO, or oxygen and NO_2. When those formed by soluble bases are decomposed by heat, they leave a strongly alkaline residue. When nitrates in the solid state are heated with *concentrated sulphuric acid*, in the presence of *copper turnings*, NO is evolved, which, in contact with the air, forms red fumes, owing to its conversion into NO_2. This experiment ought to be conducted in a narrow test-tube. The colour is best observed by looking into the test-tube lengthways. Employ a solution of potassium nitrate for this experiment.

736. *Chlorates.*—All the chlorates, as will be seen by the table, are soluble in water. They are all decomposed by heat. Those of the alkalies and alkaline earths disengage oxygen, a chloride remaining which is *neutral*, while the corresponding nitrates under the same circumstances leave a strongly alkaline residue. The other chlorates disengage by heat a mixture of oxygen and chlorine, leaving an oxide or oxychloride of the metal. The chlorates are energetic supporters of combustion, deflagrate on heated coals, and produce violent detonations when heated with very combustible bodies, such as charcoal, sulphur, and phosphorus. To detect this acid, add to a *small quantity* of the solid substance under examination a few drops of *concentrated sulphuric acid* in the cold. The chlorate will be decomposed, POTASSIUM SULPHATE and POTASSIUM PERCHLORATE ($KClO_4$) being formed, along with a GREENISH-YELLOW-COLOURED GAS, ClO_2, which escapes. The application of heat must be avoided; and the quantities operated upon should be small, to prevent any loud explosion taking place. Employ a few grains of potassium chlorate for this experiment.

737. *Orthophosphates.*—The alkaline phosphates are the only phosphates which are soluble in water; all the others are insoluble in it, but dissolve readily in an acid liquid. The phosphates evince no apparent reaction with oil of vitriol, and are thus instantly distinguished from all salts which disengage acid vapours under the same circumstances. Silver nitrate produces, in aqueous solutions of the phosphates, a

yellow precipitate of silver phosphate (Ag_3PO_4). Employ disodium hydric phosphate for this experiment.

738. *Arseniates and Arsenites.*—Sulphuretted hydrogen produces, in acidified solutions of these salts, a yellow precipitate, soluble in ammonia; the precipitate is formed immediately in solutions of the arsenites, but only after the lapse of some time in the solutions of the arseniates. Employ sodium or potassium arsenite and arseniate for this experiment. The arseniates are isomorphous with the corresponding phosphates. The two classes of salts exhibit in their general character the strongest possible similarity.

739. *Borates.*—The borates are not decomposed on ignition; they all, *even the acid salts*, manifest an alkaline reaction. Boraric acid turns turmeric paper brown, like an alkali: a borate may therefore be detected by adding hydrochloric acid to the solution of it, and then dipping a piece of turmeric paper into it, and afterwards drying the paper at a gentle heat; if it dries of a distinctly brown tint, boracic acid is present. Employ biborate of soda for this experiment.

740. *Chromates.*—These salts are isomorphous with the corresponding sulphates; all the salts are coloured. There are three series of these salts, viz., neutral, basic, and anhydro salts; but no acid salts have as yet been obtained; they decompose as soon as formed into anhydro salts and water; it is possible they may exist at a low temperature. Boiled with an excess of HCl they are decomposed, Cr_2Cl_6 being formed, which is known by the change of colour, as the solution of the chloride is green. Insoluble chromates fused with nitre yield potassium chromate, which may be dissolved out by water.

741. *Silicates.*—The silicates vary very much in chemical constitution, some correspond to H_2SiO_4, others correspond to the dibasic acid. Silicates, various in composition, occur very abundantly as natural minerals. All the silicates, excepting the alkaline silicates, are insoluble in water, and the solubility of the latter salts in water increases with increase of the alkali. These salts are known as *soluble glass* or *water-glass.* Some of the salts insoluble in water dissolve with decomposition in boiling HCl or HNO_3, the silicic acid separating as a gelatinous mass; those insoluble in these acids are decomposed on

greater or less affinity for oxygen. All the carbonates, without exception, are decomposed when heated to a very high temperature with charcoal, carbon monoxide being disengaged.

745. The alkaline carbonates and bicarbonates affect test-paper in the manner of a free alkali.

746. *Oxalates.*—The alkaline oxalates are soluble in water; the rest are almost all insoluble in that liquid, but soluble in dilute acids. They are all decomposed by heat. The oxalates of the more easily reducible metals, as Ag, Hg, Cu, give off CO_2, the metal being left in the metallic state; the oxalates of the less reducible metals, whose carbonates are decomposed by heat, as Zn, Mg, &c., give off CO_2 and CO, the metal being left as oxide. The oxalates of the metals whose carbonates are not decomposed by heat, as K and Na, or *not easily* decomposed, as Ba, Sr, and Ca, give off CO, the metal being left as carbonate. The oxalates like the acid, when heated with concentrated H_2SO_4, give off, with effervescence, CO_2 and CO, and if a light be applied to these gases, as they issue from the mouth of the test-tube, the CO will burn with a blue flame. Soluble calcium salts and lime-water produce, even in highly *dilute* aqueous solutions of oxalates, a white precipitate of calcium oxalate, which is insoluble in acetic acid.

Miscellaneous Experiments.

597. Name the metals whose solutions will give a precipitate with hydrochloric acid, and then add to a solution of each of the metals hydrochloric acid, in order to verify your statement.

598. Decompose ammonium chloride (in solution) in such a way that NH_3 will be evolved.

599. Prepare ferric hydrate from the sesquichloride.

600. Prepare some zinc oxide from the carbonate.

601. Explain by equations the chemical changes which would ensue on adding to an aqueous solution of barium chloride ammonia, and then a solution of ferrous sulphate, and afterwards perform the experiment.

602. Explain by equations the chemical changes which would ensue on adding to an aqueous solution of calcium chloride caustic soda; add then ammonium carbonate, and afterwards perform the experiment.

603. Prepare some manganese sulphide from the chloride.

604. Dissolve some soluble barium salt in water, and prepare from that solution the phosphate, chromate, carbonate, and sulphate. For this purpose divide the solution into four separate portions; add to one of the four portions a soluble phosphate, to another a soluble chromate, to a third a soluble carbonate, and to the remaining portion a soluble sulphate.

605. Prepare anhydrous oxide, hydrated oxide, and sulphide of copper, from a salt of that metal, which is soluble in water.

606. Prepare a small quantity of sulphate, carbonate, and chloride of lead, from a salt of that metal which is soluble in water.

607. Prepare some barium sulphate from the carbonate.[1]

608. Prepare some zinc carbonate from the sulphide.

609. Dissolve a barium salt and a mercurous salt, which can exist together without decomposition, in an appropriate quantity of water; throw down the mercury from the solution containing the two salts, by adding to the solution some acid, base, or salt, which will precipitate it either as an insoluble oxide or as an insoluble salt; filter off from the precipitate thus formed, and to the filtrate—which ought, if sufficient of the substance employed to precipitate the mercury were used, only to contain the barium compound (disregarding the substance employed to precipitate the mercury)—add some acid, base, or salt, which will precipitate the barium.

610. Dissolve a mercuric salt and a calcium one in water, and separate them in a similar manner.

611. Dissolve a calcium salt, a zinc salt, and a ferric salt, and separate them in a similar manner.

612. A manufacturing chemist has a quantity of impure ammonium chloride which he desires to purify; the impurity is ferric chloride. If he were to dissolve the impure ammonium salt in water, could he precipitate the iron in such a way that chloride of ammonium would be the only substance remaining in solution?

747. In each of the three following tables the action of certain reagents upon certain metals in solution is contrasted by placing them in parallel columns. The student,[2] when performing the experiment, must note the *agreements* and *differences* the three metallic solutions display with each of the three reagents. The experiments must be made in the following order; taking the first table as an example:—the experiments on all the three metallic solutions must be made with the reagent HCl before acting upon any one of them with either of the other two reagents; when these are completed, he must follow the same course with the reagent ammonia, and finally with the reagent oxalic acid. At the conclusion of each experiment he must, before proceeding to the next experiment, carefully note the change produced, if any, and explain by a diagram the decomposition produced, if any precipitation occurs. When all the experiments have been completed in any one table, before proceeding to the next table, the student

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[1] Before insoluble salts can be prepared from insoluble ones, the latter must, by an acid or other means, be brought into a state of solution.

[2] If, from want of room or other causes, the students cannot make the experiments for themselves, the teacher ought to have large diagrams of the tables suspended in the class-room, and after he has made the experiments, he ought to make each student state in writing all the different methods which could be adopted for the separation of the three bases.

must devise a method, or—if more than one could be adopted—all the methods which could be employed by the aid of the three reagents, for the detection and separation of the three metals, whether separated in the form of oxides or salts. The teacher ought then to give the student three mixed solutions of the metals to analyse, and the three solutions should not

TABLE I.

SILVER OXIDE. (Ag_2O).	FERRIC OXIDE. (Fe_2O_3).	LIME. (CaO).
1. *Hydrochloric acid* precipitates silver from its neutral and acid solutions in the form of chloride ($AgCl$), because silver chloride is insoluble in neutral and acid solutions.	1. *Hydrochloric acid* produces no precipitate in solutions of iron in the ferric state, because ferric chloride is soluble.	1. *Hydrochloric acid* does not precipitate lime from its solutions, because calcium chloride is soluble.
2. *Ammonia* (NH_4HO) precipitates silver oxide from its solutions; but an excess of ammonia redissolves it.	2. *Ammonia* precipitates ferric hydrate from ferric solutions, and the hydrate is not redissolved by an excess of ammonia.	2. *Ammonia* does not precipitate lime from its solutions.
3. *Oxalic acid* produces in neutral, but not in ammoniacal solutions, a precipitate of silver oxalate, as the oxalate is soluble in ammonia.	3. *Oxalic acid* does not precipitate iron from its ferric solutions, as ferric oxalate is soluble.	3. *Oxalic acid* precipitates lime as oxalate from its neutral and alkaline solutions.

TABLE II.

CUPRIC OXIDE. (CuO).	ALUMINA. (Al_2O_3).	BARYTA. (BaO).
1. *Hydrosulphuric acid* precipitates copper from its acid solutions as sulphide (CuS).	1. *Hydrosulphuric acid* does not precipitate alumina from its acid solutions.	1. *Hydrosulphuric acid* does not precipitate baryta from its solutions.
2. *Ammonia* precipitates copper from its acid solutions, but an excess of ammonia redissolves the precipitate.	2. *Ammonia* precipitates alumina from its solutions in the form of hydrate ($Al_2H_6O_6$), and an excess of ammonia does not redissolve it.	2. *Ammonia* does not precipitate baryta from its solutions.
3. *Sulphuric acid* produces no precipitate in solutions of copper, because cupric sulphate is soluble.	3. *Sulphuric acid* produces no precipitate in solutions of alumina, because aluminium sulphate is soluble.	3. *Sulphuric acid* precipitates baryta from its solutions, because barium sulphate is insoluble.

TABLE III.

SILVER OXIDE. (Ag_2O).	CUPRIC OXIDE. (CuO).	ZINC OXIDE. (ZnO).
1. *Hydrochloric acid* precipitates silver from its neutral and acid solutions as chloride.	1. *Hydrochloric acid* causes no precipitate in solutions of copper.	1. *Hydrochloric acid* causes no precipitate in solutions of zinc.
2. *Hydrosulphuric acid* precipitates silver from its acid solutions as sulphide (Ag_2S).	2. *Hydrosulphuric acid* precipitates copper from its acid solutions as sulphide (CuS).	2. *Hydrosulphuric acid* does not precipitate zinc from its acid solutions.
3. *Ammonium sulphide* precipitates silver from its neutral and alkaline solutions as sulphide.	3. *Ammonium sulphide* precipitates copper from its neutral and alkaline solutions as sulphide.	3. *Ammonium sulphide* precipitates zinc from its neutral and alkaline solutions as sulphide (ZnS).

contain all the three metals, as in chemical analysis it is sometimes more difficult to prove the absence of a substance than to prove the presence of one.

748. The student has now become acquainted, in a general way, with some of the methods the chemist adopts for distinguishing and separating substances from each other; the discovery of the various elements or compounds in a complex compound, or in a mixture of substances, and thus the separation of them from one another, is the object of that branch of practical chemistry termed Qualitative Analysis. There is another branch of analysis termed Quantitative Analysis. The object of this branch of analysis is to determine the quantity of each element or compound present in any substance. This branch of analysis is divided into gravimetric and volumetric analysis. We will simply explain gravimetric analysis by one illustration. Suppose we wished to estimate the quantity of barium in any substance containing it: the first thing, if it were a solid, after weighing a certain quantity of it, would be to bring it into solution, and then precipitate the barium in some insoluble form, say as sulphate, then collect on a paper filter the insoluble barium compound, free it from every other substance by washing it with distilled water by the aid of the wash-bottle (fig. 58), dry it, then weigh it, and finally calculate the amount of barium it contained.

749. Volumetric analysis is the branch of quantitative analysis which I wish more particularly to make the student acquainted with, as it is one which enables me to show him in a brief space how the knowledge he has acquired in the previous lessons can be usefully and practically applied. In volumetric analysis the quantities are determined, not by separating and weighing a definite quantity of the substance we wish to estimate, as is done in gravimetric analysis, but by adding to a solution of the substance being analysed *another* of definite strength called a *standard* solution, until the desired reaction is obtained; as, for example, adding, from a *graduated* vessel, a standard solution of sulphuric, hydrochloric, or nitric acid to a measured volume of *one* of the alkalies, until the solution has no action on *blue* or *red* litmus paper; or by adding a standard solution of one of the alkalies to a measured

quantity of one of these acids until the solution is *neutral*, that is, has no action either upon blue or red litmus paper. The determination of the amount of free acid is termed *Acidimetry*; the estimation of an alkali, whether in the free state or in combination with carbonic acid, is termed *Alkalimetry*.

750. One of the most important instruments in volumetric analysis is the burette. There are several forms of it, but in each form it consists of a graduated tube, so that the quantity of any liquid delivered from it can be accurately measured. The one represented in fig. 71 is called Binks's burette; it is, as shown, an upright, graduated tube provided with a spout for the delivery of a solution, and another opening, for the introduction of a solution and by the closing of which by the thumb the liquid is prevented from escaping from the spout, and by not closing the opening perfectly air-tight the flow of the liquid can be regulated from the other opening. This form requires a wooden stand, into which the lower end is placed when the instrument is not in use.

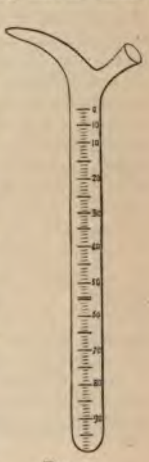

Fig. 71.

751. Graduated pipettes and graduated flasks are also required in volumetric analysis; but as we are not giving an analytical course, but only a few examples of volumetric determinations by *precipitation*, by *saturation*, by *oxidation*, by *decolourisation*, &c., in order to extend the student's knowledge of the chemical actions of substances upon each other, and some of their applications, we shall not enter into any further description of the instruments required.

Experimental Exercises in Volumetric Analysis.

613. *Estimation of Chlorine in soluble Chlorides.*—The student is already aware that silver chloride is insoluble; if, therefore, we were to add a measured volume of a standard solution of silver nitrate to a weighed quantity of a chloride, or a substance containing a chloride, which we had dissolved in water, until no precipitation took place, we could, by knowing the quantity of silver nitrate in the quantity of silver solution employed, calculate the amount of chlorine in the chloride; but there is a difficulty in ascertaining the exact point of complete precipitation; two or three drops of a not over strong solution of potassium chromate are therefore added to

the soluble chloride. Silver chromate, which is *red* in colour and *insoluble*, is not precipitated to the slightest extent until *all* the chlorine has been precipitated; as soon, then, as the faintest red precipitate *permanently* appears, we know that all the chlorine has been precipitated.

The standard solution of silver may be prepared by dissolving 17 grams of pure crystallised silver nitrate in distilled water, and adding the water until the solution *exactly* measures a litre. A solution of this strength contains 0·0108 gram. of silver in one c.c. of the solution; this quantity of silver combines with ·00355 gram. of chlorine to form the chloride.[1]

Dissolve 0·5 gram. of common salt in distilled water, add a few drops of a solution of potassium chromate, then add slowly from the burette the standard silver solution, stirring the mixture continually with a glass rod. As the precipitation approaches completion, add only a few drops of the silver solution at a time, and when the faintest permanent reddish colour is produced the operation is completed; then read off from the burette the number of c.c. of silver solution employed, and calculate the quantity of sodium chloride contained in the ·5 gram. of salt. A second determination ought always to be made; if the two results agree, the estimation is concluded to be correct; but if they do not agree, the estimation must be repeated until two estimations do agree, or within a very small fraction; this must be done not only in this, but in all the other examples.

On the addition of the silver solution a red colour will appear before all the chlorine is precipitated, but it disappears on stirring or shaking the solution; it is only when it becomes permanent that the operation is finished. As silver chromate is soluble in acids, as the table shows, the chloride solution must be *neutral* or *feebly* alkaline; if *acid* it must be neutralised by sodium carbonate *free* from a chloride.

614. *Estimation of Cyanogen in the Alkaline Cyanides.*—Silver cyanide is insoluble in water, but a molecule of this salt unites with a molecule of an alkaline cyanide, forming a double salt which is soluble, thus:

$$2KCy + AgNO_3 = AgCy,KCy + KNO_3.$$

But if more silver solution is added than is required to form the double salt, AgCy begins to precipitate. Advantage has been taken of this fact to estimate cyanogen volumetrically: a standard solution of silver nitrate is added until a *faint* but *permanent turbidity* is produced; when this occurs we know the proper quantity of silver nitrate has been added.

Weigh out about one gram. of commercial potassium cyanide, dissolve it in water, then add slowly to the solution from the burette the standard silver solution employed in the last experiment; and as it approaches completion add it drop by drop, as the AgCy, which first precipitates, dissolves with less and less rapidity in the alkaline cyanide as the complete formation of the double cyanide is approached. When the faintest permanent turbidity is produced, the number of c.c. of silver solution must be read off, and the quantity of pure KCy in the quantity of the commercial salt employed must be calculated; and afterwards a second determination ought to be made.

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[1] It is necessary when substances have to be correctly estimated, to verify by experiment that the standard solution is of the definite strength; thus, for example, after making a standard silver solution, we ascertain by experiment whether every 100 c.c. of it exactly precipitates the chlorine in every ·585 gram. of pure sodium chloride; but, as we are only illustrating the principle, it is not necessary to do so in this and most of the other examples.

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standardised by dissolving a weighed quantity of soft thin iron wire in HCl, taking care during the dissolving the solution does not come in contact with the air, as solutions of ferrous iron rapidly absorb oxygen from the air, the metal becoming converted into the ferric state. We will assume, however, that 100 c.c. of the solution corresponds exactly to ·56 gram. of Fe; and we will show how to estimate iron by employing 2·78 grams. of ferrous sulphate ($FeSO_4.7\ H_2O$).

Dissolve a few grains of potassium ferricyanide in water, and, by dipping a glass rod into the solution, place a dozen or so of separate drops of the solution on a white porcelain plate—an ordinary dinner plate will do, Dissolve 2·74 grams. of the ferrous sulphate in cold water. If the salt was perfectly free from any of the iron, being in the ferric state, which it really is, it would require 100 c.c. of the bichromate solution to preoxidise all the iron. Add some c.c. of the bichromate solution to the sulphate solution, say 60, then take a drop out by means of a glass rod and drop it on one of the ferricyanide drops; if a blue colour is produced add more of the bichromate; repeat the test, and so continue until a blue tinge just ceases to be produced; then read off the number of c.c. of bichromate solution which have been required, and calculate the amount of iron, in the ferrous state, the salt contained.

When it is desired to estimate iron which is in the ferric state by this method, it is necessary to reduce it to the ferrous state; this is generally done by adding to the acid solution of the iron pure zinc, or an alkaline sulphite, or sulphuretted hydrogen; these are the reducing agents usually employed for the purpose.

617. *Alkalimetry.*—The alkalies in the free state and in their carbonates are volumetrically determined by means of a standard acid, sulphuric being the one generally employed; and for this purpose it is best to obtain the purest acid. Measure out 30 c.c. of the acid, add to it three or four times its volume of distilled water, and allow the liquid to cool; when cold make up the solution with water to a litre.

Ignite some pure sodium carbonate; when cold, weigh out a gramme, and place it in an evaporating dish; add distilled water and heat to near the boiling-point, then add about 10 c.c. of litmus solution, and subsequently the acid solution from the burette. On the evolution of CO_2 the colour of the solution will change to *claret* or *pale red*; heat the solution to the boiling-point for a few minutes to expel the CO_2; on the expulsion the blue colour will return; then add more acid, boil again, and so continue until the solution is of a neutral tint, which is not affected by boiling the liquid. Read off the number of c.c. of acid required to effect this, and repeat the experiment.

We desire to have the acid of such a strength that 49 parts of H_2SO_4 are contained in a litre. As 98 parts of H_2SO_4 neutralise 106 parts of Na_2CO_3, one litre of the standard acid ought to neutralise 53 parts of Na_2CO_3, and 100 c.c. 5·3 parts. Supposing 95 c.c. of the acid neutralised 5·3 parts of Na_2CO_3, the acid solution contains one-twentieth too much acid; therefore, to obtain it of the strength required, every 95 measures must be diluted with five measures of water. Suppose, however, that 105 c.c. of the acid were required to neutralise 5·3 parts of Na_2CO_3, the acid in this case contains one-twentieth too much water; therefore, one-twentieth more of the strong acid requires to be added to it.

618. *Acidimetry.*—The estimation of the amount of a free acid is termed acidimetry; a standard solution of a free alkali, and a solution of litmus, are the agents required.

Pure caustic soda prepared from metallic sodium is the alkali best suited for most acidimetrical estimations; it can be obtained from chemists who sell laboratory reagents. Dissolve 40 grams. of the NaHO in distilled water and make up to a litre. Ascertain whether the NaHO solution is of the required strength by a standard acid; if it is not, it can be correctly adjusted in the manner described under Alkalimetry. The solution of the alkali is added from the burette to a *known quantity* of the acid we wish to determine until the solution is neutral—that is, neither changes the colour of blue or red litmus.

Dissolve one gram. of commercial oxalic acid in distilled water, add some solution of litmus, then add from the burette the solution of the alkali until the solution is neutral. Read off from the burette the number of c.c. of the soda solution employed, then calculate the quantity of acid, and afterwards repeat the experiment.

Solutions of the caustic alkalies, as they rapidly absorb CO_2 from the air, should be carefully preserved in stoppered bottles.

619. The carbonates of the alkaline earths, although insoluble in water, can be estimated volumetrically; for this purpose, we add to a known weight of one of these carbonates a measured quantity of a standard solution of hydrochloric acid (sulphuric acid cannot be conveniently employed because $BaSO_4$ and $SrSO_4$ are insoluble, and $CaSO_4$ only slightly soluble in water) sufficient to render the litmus which has been added to the solution distinctly red; then warm the solution, but do not actually boil it, until all the CO_2 is expelled; then add to the solution, which ought still to turn blue litmus red, from the burette a standard solution of one of the alkalies until the solution is neutral to test paper. The amount of acid which remains free having been ascertained, we arrive at the quantity the carbonate neutralised; we thus ascertain the quantity of the particular carbonate present in the substance.

Dissolve 0·5 gram. of marble in a measured quantity of a standard solution of HCl, warm the solution, then add from the burette a standard solution of one of the alkalies, then calculate the amount of $CaCO_3$ in the 0·5 gram., and repeat the experiment.

752. In concluding this outline of volumetric methods we must inform the student what is meant by a *normal* standard solution; for he will meet in books which treat on this method of analysis the terms *normal* solution, *quantinormal*, *decinormal*, and *centinormal.* Normal solutions are so prepared that one litre at 16° C. contains the hydrogen equivalent of the active reagent weighed in grams. $(H=1)$; a decinormal solution is one-tenth the strength of a normal one, &c.

753. Standard solutions can be purchased at the operative chemists'; some teachers may find it more convenient to purchase them than to prepare them.

Examination Questions.

620. To what substance is the hardness of water due? Describe a process by which it can be softened.

621. Name the sulphates which suffer no change on being heated.

	Substances added.	*Substances formed or set free.*
648.	Sol. of potassium chromate.	Barium chromate.
	Sol. of barium chloride.	And ——— ?
649.	Sol. of sodium hydrate.	Manganous hydrate.
	Sol. of manganous chloride.	And ——— ?
650.	Sol. of ammonium carbonate.	Calcium carbonate.
	Sol. of calcium nitrate.	And ——— ?
651.	Hydrosulphuric acid.	Plumbic sulphide.
	Sol. of plumbic nitrate.	And ——— ?
652.	Sol. of sodium sulphate.	Barium sulphate.
	Sol. of barium chloride.	And ——— ?
653.	Sol. of oxalic acid.	Barium oxalate.
	Sol. of barium nitrate.	And ——— ?
654.	Ammonium sulphide.	Cobaltous sulphide.
	Sol. of cobaltous nitrate.	And ——— ?
655.	Hydrochloric acid.	Mercurous chloride.
	Sol. of mercurous nitrate.	And ——— ?
656.	Hydrosulphuric acid.	Mercuric sulphide.
	Sol. of mercuric chloride.	And ——— ?

657. When potassium chlorate in the solid state is ignited, it is decomposed into oxygen and potassium chloride. How many atoms of oxygen will be liberated from one molecule of the salt?

658. When ammonium nitrate in the solid state is heated, it is decomposed into nitrogen monoxide (laughing gas, N_2O) and some other substance. What is the other substance, and how many molecules of each will be formed from one molecule of the salt?

659. We know the substances that are formed on heating NH_4NO_3; what substances will be formed on igniting ammonium nitrite, NH_4NO_2; and how many molecules of the substances formed will be produced from one molecule of the salt?

	Substances added.	*Substances formed or set free.*
660.	Hydrochloric acid.	Ferric chloride.
	Ferric oxide.	And ——— ?
661.	Sol. of cupric chloride.	Cuprous iodide (Cu_2I_2).
	Sol. of potassium iodide.	And ——— ?
662.	Sol. of hydric disodium arsenite.	Hydric cupric arsenite.
	Sol. of cupric sulphate.	And ——— ?
663.	Sol. of sodium hydrate.	Chromic hydrate.
	Sol. of chromic chloride.	And ——— ?
664.	Sulphuric acid.	Manganous sulphate.
	Manganese dioxide.	And ——— ?
665.	Antimonious chloride.	Oxychloride of antimony ($SbCl_3,Sb_2O_3$).
	Water.	And ——— ?
666.	Sol. of ferric sulphate.	Ferrous sulphate.
	Sulphurous acid.	And ——— ?
667.	Sol. of sodium hydrate.	Aluminium hydrate.
	Sol. of aluminium sulphate.	And ——— ?
668.	Sol. of hydric disodium phosphate.	Hydric barium phosphate.
	Sol. of barium nitrate.	And ——— ?

	Substances added.	*Substances formed or set free.*
669.	Sol. of hydric disodium phosphate. Sol. of argentic nitrate.	Triargentic phosphate. And ——— ?
670.	Sol. of hydric disodium phosphate. Sol. of cupric sulphate.	Hydric cupric phosphate. And ——— ?
671.	Hydrosulphuric acid. Sol. of ferric sulphate.	Ferrous sulphate. And ——— ?
672.	Arsenous acid. Sulphuric acid. Zinc.	Arsine. And ——— ?
673.	Hydrochloric acid. Manganic dioxide.	Manganous chloride. And ——— ?
674.	Sol. of hydric disodium phosphate. Sol. of ammonia. Sol. of magnesium sulphate.	Ammonium magnesium phosphate. And ——— ?
675.	Sol. of cupric nitrate. Sol. of ferrous sulphate. Sol. of potassium iodide.	Cuprous iodide. Ferric sulphate. And ——— ?
676.	Sol. of aluminium sulphate. Sol. of ammonium carbonate.	Aluminium hydrate. And ——— ?
677.	Sol. of chromic chloride. Sulphide of ammonium.	Chromic hydrate. And ——— ?
678.	Sol. of ferric chloride. Sol. of ammonium carbonate.	Ferric hydrate. And ——— ?

679. When potassium bichromate and sulphuric acid are heated together, chromic sulphate and other substances are formed. What are the other substances?

680. Chlorine converts solutions of ferrous salts into ferric salts; show the changes which take place when chlorine is added to a solution of ferrous sulphate.

	Substances added.	*Substances formed or set free.*
681.	Sol. of ferrous chloride. Sol. of potassium bichromate. Hydrochloric acid.	Ferric chloride. And ——— ?
682.	Manganic dioxide. Sodium chloride. Sulphuric acid.	Manganous sulphate. And ——— ?
683.	Sol. of potassium hydrate. Potassium nitrate. Zinc.	Potassium zinc oxide. And ——— ?
684.	Boiling sol. of potassium hydrate. Sol. arsenious acid. Zinc.	Potassium zinc oxide. And ——— ?
685.	Antimonious sulphide. Hydrochloric acid.	Antimonious chloride. And ——— ?
686.	Sulphurous acid. Sol. of iodine.	Sulphuric acid. And ——— ?
687.	Sol. sodium thiosulphate ($Na_2S_2O_3$). Sol. of iodine.	Sodium tetrathionate ($Na_2S_4O_6$). And ——— ?
688.	Sol. of oxalic acid. Manganese dioxide. Sulphuric acid.	Manganous sulphate. And ——— ?

	Substances added.	*Substances formed or set free.*
689.	Cuprous oxide (Cu_2O). Sol. of ferric chloride. Hydrochloric acid.	Ferrous chloride. And ——— ?
690.	Ferrous sulphate. Sol. of potassium permanganate ($K_2Mn_2O_8$). Sulphuric acid.	Ferric sulphate. And ——— ?
691.	Sol. of potassium chlorate. Ferrous chloride. Hydrochloric acid.	Ferric chloride. And ——— ?
692.	Sol. of potassium iodide. Manganese dioxide. Sulphuric acid.	Manganous sulphate. And ——— ?
693.	Arsenious zincide (As_2Zn_3). Sulphuric acid.	Zinc sulphate. And ——— ?

694. Name two acids, either of which, on being added to zinc, would yield hydrogen, and show the reactions by means of equations.

695. Give as many methods as you can, from the preceding exercises in this chapter, for the preparation of oxygen.

696. When ammonium nitrate is heated, what are the products of decomposition ?

697. Name one or more substances which reduce ferric salts to the state of ferrous salts.

698. Explain the following chemical reactions by means of symbols :

a. The decomposition of steam by zinc.

b. The decomposition of water by potassium.

c. The formation of hydrochloric acid and sodium sulphate from common salt and sulphuric acid.

d. The action of chlorine on sulphuretted hydrogen.

e. The formation of sulphuretted hydrogen from ferrous sulphide and sulphuric acid.

699. What substance would you add to common salt and sulphuric acid if you wished to obtain chlorine.

700. What substance would you employ if you wished to precipitate by copper sulphate the whole of the iodine from potassium iodide in the form of cuprous iodide ?

701. How would you prepare chlorine ?

702. When a small quantity of water is added to pentachloride of phosphorus, oxytrichloride of phosphorus ($POCl_3$) and some other substance are formed. When an excess of water is added to the pentachloride, phosphoric acid and some other substance are formed. Express each reaction by an equation.

703. When manganese dioxide is strongly heated, it is decomposed ; a complex oxide of manganese, Mn_3O_4, is produced. What other substance is formed or set free ?

704. When silicic anhydride, calcic fluoride, and sulphuric acid are heated together, calcium sulphate and some other substances are formed. What are the other substances, and how many molecules of each are formed ?

705. If sulphurous anhydride be transmitted through water in which finely divided manganese dioxide is suspended, manganous hyposulphate (manganous dithionate), $Mn''S_2O_6$, is formed if the liquid be kept cool. Express the reaction by an equation.

706. Silicon tetrafluoride is decomposed by water into hydrofluosilicic acid, H_2SiF_6, and some other substance. What is the other substance, and how many molecules of each are formed?

707. If manganous sulphide is digested in a solution of ammonium sulphate, ammonium trithionate $(NH_4)_2S_3O_6$, and other substances are formed. What are the other substances, and show by an equation how many molecules of each are produced?

708. If chlorine be transmitted through a weak and cold solution of potassium hydrate, potassium hypochlorite ($KClO$) and other substances are formed. Express the reaction by an equation.

709. If chlorine be transmitted through a hot or concentrated solution of potassium hydrate, potassium chlorate and other substances are formed. Express the reaction by an equation.

710. If a current of hydrogen sulphide be transmitted through a solution of sulphurous acid, pentathionic acid ($H_2S_5O_6$) and other substances are formed. Express the reaction by an equation.

711. When calcium hydrate and an excess of sulphur are boiled together in water, pentasulphide of calcium and other substances are formed. Express the reaction by an equation.

712. When phosphorus is boiled with barium hydrate in water, barium hypophosphite ($Ba2PH_2O_2$) and another substance are formed. What is the other substance?

713. When a solution of ammonia is added to a solution of mercuric chloride, dichloride of dimercurammonium ($Hg''_2H_4N_2Cl_2$) and another substance are formed. Express the reaction by an equation.

714. If phosphine is passed through a solution of silver nitrate, phosphoric acid and other substances are formed. Express the reaction by an equation.

715. Can you name any method for the production of ammonia from nitric acid?

716. What is the action of chlorine on a dilute solution of potash?

717. What is the composition of arsine? and state as many methods as you can for its preparation.

718. Name one or more methods for the preparation of hydrogen sulphide.

719. What is the action of ammonia on mercuric chloride?

720. Give the action of potassium permanganate on a solution of a ferrous salt in the presence of sulphuric or hydrochloric acid.

721. Give a method for the preparation of potassium chlorate.

722. How would you prepare phosphine?

723. Sodium thiosulphate is used as an antichlore for removing the last traces of chlorine from substances bleached by that agent. Show by an equation how it acts as an antichlore.

724. When an iodide or a bromide is treated with MnO_2 and H_2SO_4, the iodine or bromine is set free. Express the reaction by an equation.

725. How would you prepare barium sulphate from a soluble barium salt?

726. Are you acquainted with any method for the preparation of an oxychloride? [1]

Appendix.

Laurent, in his 'Chemical Method,' recommends that when substitutional reactions are represented in the form of an equation, the second member of the equation should be written underneath the first. This is a very good plan in many cases, for we at once perceive, when the equation is written in this manner, that to each term of the superior member there is a corresponding term in the inferior one, as shown in the following examples :

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[1] *The student must now return to Chap. VI. and complete the study of it, commencing at par.* 362.

1st.—Action of chlorine upon hydrogen . . . $\left\{\begin{array}{l}\mathrm{Cl\ Cl + HH =}\\ \mathrm{H\ \ Cl + H\,Cl.}\end{array}\right.$

2nd.—Chlorine upon potassium hydrate . . $\left\{\begin{array}{l}\mathrm{Cl\ Cl + KHO + KHO =}\\ \mathrm{K\ \ Cl + K\,ClO + HHO.}\end{array}\right.$

3rd.—Chlorine upon potassium hydrate . . $\left\{\begin{array}{l}\mathrm{3\ Cl\ Cl + 2\ KHO + KHO\quad + 3\ KHO =}\\ \mathrm{3\ K\ \ Cl + 2\ KCl\quad + KClO_3\ + 3\ HHO.}\end{array}\right.$

4th.—Action of chlorine and iodine on water . $\left\{\begin{array}{l}\mathrm{5\ Cl\ Cl + 5\ HHO + HHO\ \ + I\,I =}\\ \mathrm{5\ H\ \ Cl + 5\ HCl\quad + H\,I\,O_3 + HIO_3.}\end{array}\right.$

5th.—Sulphuric acid upon dioxide of manganese . $\left\{\begin{array}{l}\mathrm{H_2SO_4\quad + MnO_2 =}\\ \mathrm{Mn_2SO_4\ + H_2O + O.}\end{array}\right.$

6th.—Sulphuric acid upon potassium carbonate . $\left\{\begin{array}{l}\mathrm{H_2SO_4 + K_2CO_3 =}\\ \mathrm{K_2SO_4 + CO_2 + H_2O.}\end{array}\right.$

In the fifth reaction, peroxide of hydrogen ought to be formed; but this body, not being capable of existing in the presence of binoxide of manganese, undergoes decomposition, and sets free its oxygen, whence the interference with the symmetry of the reaction.

In the sixth reaction, the symmetry is destroyed, because H_2CO_3 being, under ordinary circumstances, incapable of existing, becomes decomposed immediately into H_2O and CO_2.

CHAPTER XV.

COMBINATION OF SIMILAR ATOMS—THE NASCENT STATE—ALLOTROPY—DETERMINATION OF THE ATOMIC WEIGHTS OF THE ELEMENTS—PHYSICAL AND CHEMICAL RELATIONS OF ATOMIC WEIGHTS—QUESTIONS.

758. WE have now to lay before the student the chemical evidence which goes to prove that the atoms of most of the elements in the free state are *united* and not *separated.* The evidence in support of this we will arrange in three classes.

759. *1st Class.*—In a great many of the chemical combinations, in which an element is one of the combining bodies, an even number of atoms of the element is required. The atoms of the element appear, therefore, to be associated in binary groups.[1] Examples :—

1. When chlorine acts upon potash, potassium hypo-

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[1] The *perissad* elements are considered only to unite in *pairs*; but the *artiad* elements may unite in groups of *two*, *three*, or *more*.

chlorite or potassium chlorate is formed. The following is the reaction when potassium hypochlorite is produced :

$$2KHO + Cl_2 = \underset{\text{Potassium hypochlorite.}}{KClO} + KCl + H_2O.$$

When potassium chlorate is formed the following reaction takes place :

$$6KHO + 3Cl_2 = 5KCl + KClO_3 + 3H_2O.$$

2. When sulphur is fused at a gentle heat with an alkaline hydrate, or boiled with an aqueous solution of the alkali, two compounds, dipotassium pentasulphide and potassium hyposulphite, are formed :

$$6KHO + 6S_2 = K_2S_2O_3 + 2K_2S_5 + 3H_2O.$$

760. When metallic sulphides oxidise in the air, both elements enter into combination with oxygen ; thus,

$$NiS + 2OO = NiSO_4$$

761. In organic acids the whole or part of the hydrogen in the acid radical can be replaced by the halogen elements, especially by chlorine, and there does not exist a single organic compound upon which we can effect any reaction by employing an *uneven* number of atoms of chlorine. In every case it is Cl_2, or a multiple thereof, which determines the reaction ; *the compound may contain an uneven number* of atoms. The following are examples of this statement :—

$$\underset{\text{Acetic acid}}{HC_2H_3O_2} + Cl_2 = \underset{\text{Monochloracetic acid}}{HC_2H_2ClO_2} + HCl.^{1}$$

$$HC_2H_3O_2 + 3Cl_2 = \underset{\text{Trichloracetic acid}}{HC_2Cl_3O_2} + 3HCl.$$

2nd Class.—There are a great number of decompositions which cannot be explained in any satisfactory manner unless we admit that *the atoms of the same elements are capable of entering into combination with each other.*

762. When a solution of potassium dichromate is poured into a strong acid solution of barium peroxide (BaO_2) in

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[1] The hydrogen in the acid radical is not replaceable by metals, but is replaceable by the halogen elements, as exemplified by the action of chlorine on acetic acid.

hydrochloric or nitric acid, a violent effervescence takes place, due to the escape of oxygen. Apart these solutions are perfectly stable, but when brought together they are both decomposed; oxygen is liberated from both compounds, the metals becoming converted into chlorides; thus—

$$10HCl+K_2CrO_4.\ CrO_3+BaO_2=Cr_2Cl_6+2KCl+BaCl_2 \\ +5H_2O+2OO.$$

The late Sir B. Brodie regarded the oxygen itself in this experiment as the true reducing agent; he believed that the chromic acid is decomposed by the oxygen of the peroxide of hydrogen, according to the same law of decomposition, and for the very same reason as it would be by hydrogen itself if a piece of zinc were thrown into the solution, the reason being the polarity of the particles induced by chemical change.

763. Before passing on to the 2nd example in the 2nd class, it is desirable to inform the student that the reason assigned for the chemical union of the atoms of the same element is the same as that assigned for the union of the atoms of different elements, and of the molecular atoms of compound substances; it is this, that at the moment of *chemical change* a chemical difference exists between the substances, whether elementary or compound, that are set *free* and *combine*, which is expressed by the terms *positive* and *negative*; therefore, when chemical combination takes place between the atoms of the *same* or *unlike* substances, a chemical difference exists between the atoms as they are to one another in a *positive* and *negative* relation.

2. Cuprous hydride, CuH, in contact with strong hydrochloric acid, is decomposed along with the acid, cuprous chloride, CuCl, being formed, and hydrogen evolved; thus—

$$\overset{+}{Cu}\overset{-}{H}+\overset{+}{H}\overset{-}{Cl}=CuCl+2HH.$$

This is perfectly analogous to the following decomposition:—

$$\overset{+}{Cu_2}\overset{-}{O}+\overset{+}{H}\overset{-}{Cl}=Cu_2Cl_2+H_2O.$$

Hydrochloric acid *scarcely* acts upon copper; the combination of hydrogen with copper, far from favouring the action of the acid upon the metal, ought, according to the idea that the atoms of the same element do not unite but remain free, to add a new obstacle to it. Hydrochloric acid also dissolves an alloy of copper and zinc, although on the copper itself it has little action. The hydride of copper is itself, in its chemical relations, an alloy; the action of the acid consequently is the same upon the compound of copper and hydrogen as upon the compound of copper and zinc, and the explanation of these facts involves similar phenomena. Thus—

$$2\overset{+}{\mathrm{H}}\overset{-}{\mathrm{Cl}}+\overset{+}{\mathrm{Zn}}\overset{-}{\mathrm{Cu}}+{}_2\overset{+}{\mathrm{Cl}}\overset{-}{\mathrm{H}}=\mathrm{ZnCl_2}+\mathrm{CuCl_2}+2\mathrm{HH}.$$

The oxides of gold, silver, and mercury, on being placed in a solution of hydrogen dioxide (H_2O_2), decompose and are decomposed by that body, although *apart* the metallic oxides and the dioxide are perfectly stable; the metallic oxides lose the whole of their oxygen, whilst the dioxide loses one-half a molecule of oxygen, a molecule of the metal and water being formed. Example:—

$$\overset{+}{\mathrm{Ag_2}}\overset{-}{\mathrm{O}}+\overset{+}{\mathrm{H_2}}\overset{-}{\mathrm{O}}\overset{+}{\mathrm{O}}=\mathrm{Ag_2}+\mathrm{OO}+\mathrm{H_2O}.$$

764. *3rd Class.*—As a chemical compound possesses properties perfectly different from those of its constituents, it follows that, if two or more atoms of the same element are capable of entering into combination, the homogeneous compound[1] and the elementary atoms in their free state will be distinguished by a difference in their properties. That the elements do not possess under all circumstances the same physical and chemical properties, but that the same element differs as much in properties under different conditions as some heterogeneous compounds differ from *their* constituents, is evidenced by the *allotropic* modifications of some of the elements,[2] and by the *nascent*, compared with the *ordinary*, *state* of the elements.

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[1] We shall call compounds consisting of a combination of the atoms of the same element, *homogeneous compounds*, to distinguish them from compounds composed of unlike elements: the latter class of compounds, composed of unlike elements, we shall call *heterogeneous compounds*.

[2] 'There appear to be four different conditions in which solid bodies may exist. They may be—1st. *Crystalline*, as diamond, garnet, felspars. 2nd. *Vitreous*, or glassy, as glass itself, transparent arsenious acid, and barley sugar.

1. The affinities of the elements in *their nascent state*[1] are much more powerful than in the ordinary state of the elements; nitrogen, for instance, cannot, except with very great difficulty, be made to unite directly with any element, but requires for its combination that one or both the elements should be in the nascent state. In order to effect the combination of chlorine and hydrogen, oxygen and hydrogen, independently of any agent, such as heat, it is necessary to bring them together in their nascent state. To effect the union of a solid with a gaseous body, it is necessary, in a number of cases, to bring them together in their nascent state; thus, no chemical combination will take place if a stream of hydrogen be passed into sulphur. If, however, the two elements are set free from other combinations at the same moment and in presence of one another, they combine together and form hydrosulphuric acid. This is also the case with hydrogen and the following substances:—phosphorus, carbon, arsenic, antimony, &c. We cannot well explain why the affinities of the elements should be so much greater at the *moment* they are set free from their (heterogeneous) combinations than they are afterwards, unless we admit that the atoms of the same elements are capable of entering into combination. If we admit the existence of molecular atoms, then there is no difficulty in accounting for the great disparity in the affinities of the elements in the two states; for in the nascent state the atoms of the element are *uncombined*, in the ordinary state they are united; consequently when two elements are set free from other combinations at the same moment and in presence of each other, the two elements enter into combination, because the *unlike* atoms have a greater affinity for each other than the *like* atoms have; or, as it would be expressed on Brodie's theory, the unlike atoms are in a greater positive and negative relation.

2. Phosphorus, sulphur, carbon, and silicon are very striking examples of the allotropic modifications or conditions of

3rd. *Amorphous*, or destitute of crystalline form altogether, as tinder, chalk, or clay. 4th. *Organised*, or arranged in masses, consisting of cells, fibres, or membranes, like the tissues of animals or vegetables, as hair, muscle, skin, wood, bark, leaves, &c.'—*Miller*.

[1] The moment elements are set free from their combinations they are said to be in their nascent state.

matter. Phosphorus can be obtained in the crystalline, vitreous, and amorphous state. In its ordinary state, it is a soft, semi-transparent, colourless, waxy-looking solid, which fumes in the air, emitting white vapours of an alliaceous odour. Its specific gravity is 1·83. It fuses at 111·5° F. It is very inflammable; it takes fire in the open air a few degrees above its fusing point. The crystalline form is equally inflammable with the vitreous. Amorphous phosphorus is red. In this state it is insoluble in a great many liquids in which the crystalline and waxy forms dissolve. The density of red phosphorus is 2·14. It has no smell; and it may be heated in the open air, without change, until the temperature reaches 500° F.: at this heat it melts, and bursts into flame. One form of sulphur has a density of 2·05 and melts at 239° F.; another form of the element has a density of 1·98, and melts at 248° F.; and a third form has a density of 1·957. These three varieties of sulphur are also distinguished from each other by a difference in some of their other properties. The student has only to contrast the properties of the diamond, graphite (the black-lead of pencils), and ordinary charcoal, to see that the properties of carbon are as various as the conditions in which it can exist. Silicon can be obtained in two distinct modifications, viz. the amorphous and the crystalline modification; and its properties in the two different states differ remarkably.

765. Allotropy is as yet unexplainable; it is probably due to the molecules of the different modifications containing a different number of atoms, as we have seen is the case with ozone, the allotropic modification of oxygen.

766. We shall conclude this, the last chapter of this volume with a description of some of the methods adopted for determining the numerical values of the atoms of the elementary bodies.

767. According to Dalton's atomic theory the atoms of the same element are all of equal size and weight; and the atoms of one element differ from those of the other elements in weight and in chemical properties; and as hydrogen is the lightest substance in nature, and the proportion in which it enters into combination is less than that of any other element, Dalton chose it as the *unit* in his scale of atomic weights, and

it is accepted as the unit wherever chemistry is studied. Dalton also assumed that the most stable compound of two elements must be binary, *i.e.* composed of one atom of each of the elements. Now, as the most stable combination of oxygen and hydrogen is water, and as in that compound the two elements are united in the proportion by weight of one part of H to eight of O, Dalton assumed the atomic weight of oxygen to be eight; the formula of the compound consequently became HO. Chemists at the present time regard the atomic weight of oxygen as double that number, viz. 16; and they have doubled the atomic weights of several of the other elements; for the atomic weights are now arrived at from more considerations than guided Dalton in fixing his scale of atomic weights. The considerations which guide chemists at the present time in arriving at the atomic weights of bodies 'are often somewhat complicated, based partly on the law of gaseous volumes, partly on the experimental results upon the specific heat of the bodies under investigation, partly on the isomorphism or similarity in crystalline form of bodies of analogous constitution; and the principle of analogy frequently enables us to determine satisfactorily the molecules and atomic weights of bodies with which we are comparatively but little acquainted.'

768. We will commence the examination of the foundation on which our present system of atomic weights and our notation rest, with the four typical compounds we employed to illustrate the volume-proportions in which gaseous and vapourisable substances unite; the compounds are HCl, H_2O, H_3N, and H_4C. The first question it would be natural for the student to ask would be this,—Why are the formulæ for water, nitrine, and methane not as simple as the one for hydrochloric acid? If these three compounds were written thus, HO, HN, and HC, then the atomic weight of oxygen would have to be reduced to 8, the atomic weight of nitrogen to 4·67, and the atomic weight of carbon to 3. In placing before the student the considerations which guide chemists in fixing the atomic weights of bodies, we will first direct his attention to the chemical action of some simple and compound radicals on these four typical compounds.

769. When sodium, for example, acts upon hydrochloric acid, the *whole* of the H in the compound and *never a portion*

of it, is always expelled ; a compound is therefore never formed containing both hydrogen, sodium, and chlorine, but it always consists solely of sodium and chlorine. If, on the other hand, we act on water with sodium, a compound consisting of sodium, oxygen, and half the hydrogen the water contained, is obtained ; the other half of the hydrogen having been expelled by the sodium, the latter taking its place. If the NaHO is heated with a further quantity of Na, the other half of the H is expelled by the Na, the compound Na_2O being formed. One half the hydrogen in water can, therefore, unlike the H in HCl, be displaced by Na. Now, if water had the same atomic constitution as HCl, sodium ought to expel the whole of the H, as it does in that compound, and never in part, because if it expelled a part, the atom of H must divide, which is contrary to the fundamental notion of atoms, namely that they are indivisible.

770. In nitrine the hydrogen is replaceable by thirds, and not, as in water, by half the amount. In methane we can displace the hydrogen by other radicals by one-fourth, or two-fourths, or three-fourths, or four-fourths.

771. By these reactions we find,—that the hydrogen in hydrochloric acid is not divisible; that in water it is divisible into two equal parts ; in nitrine into three parts, and not into two parts as in water, nor into four parts as in methane ; in methane it is divisible into four parts, and not into three parts as in nitrine. The student will therefore see that these reactions afford strong evidence for the belief that a molecule of hydrochloric acid contains one atom of H ; that a molecule of water contains two atoms of hydrogen ; that a molecule of nitrine contains three atoms of that element ; and that a molecule of methane contains four atoms of it.

772. We further find by chemical reactions that when water is the agent or resultant of a reaction it must contain two atoms or some multiple of two atoms, of hydrogen ; in like manner when nitrine is the agent or resultant it cannot be expressed with less than three atoms of that element ; and when methane is the agent or resultant it cannot be expressed with less than four atoms of it.

773. We have learned in studying the duality of the elementary molecules that 35·5 parts by weight of chlorine can replace one part of hydrogen ; it is therefore accepted, from this and other evidence, that 35·5 is the relative weight of an atom of chlorine. One unit weight of hydrogen is never replaced by 8 of oxygen ; oxygen always replaces hydrogen in the proportion of 16 parts to 2 of hydrogen ; the relative weight of an atom of oxygen is taken to be 16. Nitrogen replaces hydrogen in the proportion of 14 of nitrogen to 3 of the latter element, 14 is therefore accepted as the relative atomic weight of nitrogen. It has been shown that 12 parts of carbon replace 4 parts of hydrogen in 3 molecules of nitrine in the formation of an organic compound called guanidine. These are a very few examples of the aid derived from the valency of the atoms in determining the atomic weights of the elements.

774. When the specific gravity of hydrogen is taken as unity the specific gravities, as we have already learned, of most of the elementary gases and vapours are expressed by numbers identical with those that have been selected to express their atomic weights. We have also learned, 'that the molecules of compounds which are gaseous or volatile without decomposition occupy two volumes, if an atom of hydrogen occupies one volume. The result of this is that the molecular weights of volatile compounds are accurately given by their densities. And if we refer these densities to that of hydrogen taken as unity, we have only to multiply the numbers obtained by 2 to find the weight of the molecules compared with that of an atom of hydrogen=1. This is a general rule.

775. 'The atomic weights of a certain number of non-metals and metals may be calculated from the molecular weights. Thus the atomic weights of phosphorus, arsenic, antimony, silicon, tin, mercury, and lead may be calculated from the molecular weights of their volatile compounds. For example—

776. 'The molecular weight of chloride of silicon (obtained by doubling its vapour-density) is 170, and analysis shows that 170 parts of chloride of silicon contain $142 = 4 \times 35·5$ of chlorine and 28 of silicon. The vapour-density and analysis of chloride of silicon assign, therefore, to this body the formula $SiCl_4$, and to silicon the atomic weight 28, for we have reason for the belief that the molecule of chloride of silicon only contains a single atom of silicon.

777. 'This example shows the use which may be made of the law of Avogadro in the determination of molecular weights and in settling atomic weights.'

778. Analogy is also, as we have stated, a guide in determining the atomic weights of the elements; because when compounds are analogous, as the selenium and tellurium compounds with those of sulphur, they are assigned similar formulæ; therefore, for example, if we assign the formula H_2S to sulphuretted hydrogen, the formula for selenuretted hydrogen must be H_2Se, and for telluretted hydrogen H_2Te; hence, from these hydrogen compounds we can determine the atomic weights of Se and Te.

779. Isomorphism is also of assistance in determining, as we have seen in the case of alumina, the molecule of a compound and the atomic weights of its constituents: for, if the composition of one substance is known, the composition of another substance, isomorphous with it, can be inferred, and the atomic weight of its constituents determined. The student must, however, bear in mind that substances having the same crystalline form are not isomorphous, unless they are able to replace each other in the same crystal.

780. The specific heats of the elementary bodies in the solid state, if equal weights of the elements are employed, differ very considerably; but if, instead of employing equal weights, quantities in proportion to their atomic weights are employed, the specific heats of the elements, with some slight exceptions, are equal. So that an atom of any simple substance, whether its atomic weight is small, as that of lithium, or large, like those of bismuth and lead, has the same capacity for heat, and requires the same quantity of heat to raise its temperature, as an atom of any other element; this quantity is called its atomic heat, and can be arrived at by multiplying the atomic weight of the element by its specific heat. Conversely, the chemist is guided in fixing the atomic weight of an element by the number he obtains by dividing the mean atomic heat (6·4) of solid elementary bodies by the specific heat of the element in the solid state. This important law was discovered by Dulong and Petit. We

have very briefly explained this law; but we shall explain it more fully in the second volume of this work, and also the *periodic law*, first discovered by Newlands.

781. We will summarise what has been set forth in the words of the late Professor Wurtz:—

(1) The present system of atomic weights is founded on the law of volumes, and is in harmony with the hypothesis of Avogadro.

(2) It is in harmony with the law of isomorphism;

(3) And it is in harmony with the law of Dulong and Petit.

Examination Questions.

727. What multipliers must be used to transform the atomic weights of the following elements into their respective molecular weights: oxygen, sulphur, zinc, phosphorus, mercury, ozone, arsenic, cadmium, and hydrogen?

728. What products are formed on burning cyanogen in oxygen; and how much oxygen is required for the combustion of one cubic inch of cyanogen? What would be the volume of the resultant gases at the initial temperature and pressure?

729. What facts and considerations can you adduce in support of the theory of duality of certain elementary molecules?

Ex. 94. Up to 260° C. air and mercurial thermometers agree, but above that point mercury expands relatively more than air; the air thermometer is above this temperature regarded as the standard instrument: see also par. 88.

Ex. 95, pars. 105, 106.

Ex. 96. The latent heat of water is between 79° C. and 80° C.; we have taken it at 79° C.; at this temperature the 11 lbs. of water will have consequently a temperature of 11° C.

Ex. 97, Expt. 74. Ex. 98, pars. 221, &c. Ex. 99, 22·6 c.c.

Ex. 100. The latent heat of water in this question is taken as 80° C.

Ex. 101. Bodies fall to the ground in consequence of the earth's attraction on *each* of their molecules; therefore, all bodies, whatever their size, ought to fall with equal rapidity, and they do fall with equal rapidity in a vacuum, but not, as we all know, in air, owing to the unequal resistance air opposes to their descent.

Ex. 102. One cubic foot, 1563·43 cubic inches.

Ex. 103. Because descending as well as ascending currents of smoke and air would take place.

Ex. 104. The air which passes up the chimney carries off the rest.

Ex. 105. The draught of a chimney is proportional to its height; hence chimneys of a shorter height are much more liable to smoke than chimneys of greater height.

Ex. 106. The volume of a gas is inversely as the pressure; the volume being decreased one-half if the pressure be doubled. Heat, on the other hand, we have learned, expands gases; see also par. 221.

Ex. 107, 534·16 c.c. Ex. 108, 273° C. Ex. 109, 5·051 litres.

CHAPTER IV.

Ex. 110, 18·08 cub. inches. Ex. 111, 18·03 cub. inches.
Ex. 112, 15·49 cub. inches. Ex. 113, 81·395 cub. inches.
Ex. 114, 637° C. Ex. 115, 13·598 cub. inches.
Ex. 116, 24·113 cub. inches. Ex. 117, 14·6 cub. inches.
Ex. 118, 11·195 cub. inches. Ex. 119, 72·77 mm.
Ex. 120, 71·24 mm. Ex. 121, 63·88 mm. Ex. 122, par. 235.
Ex. 123, pars. 237, 238, 241, 248. Ex. 124, par. 281.

Ex. 125. On account of the greater pressure.

Ex. 126. If they were filled, they would, on ascending, tend to burst, owing to the gas expanding by reason of the pressure of the atmosphere diminishing, pars. 194, 195. They cease to ascend, and can only take a horizontal direction when the buoyancy becomes equal to the weight of the balloon.

Ex. 127. Moisture is always present in air, and the quantity varies greatly; warm air can take up more than cold air, par. 285. And if a confined portion of air be heated, and no additional moisture be added, we experience in such an atmosphere unpleasant sensations, owing to the rapidity with which evaporation takes place; hence the skin feels dry; and it is to remedy this that a plate containing water is frequently placed on the top of stoves.

Ex. 128. In one case it would be converted into the gaseous state, and in the other case it would be solidified.

Ex. 129, par. 239. Ex. 130, par. 255.

Ex. 131. The dew-point is a name used to denote the lowest temperature to which the air can be cooled down without depositing some of its moisture.

Ex. 132. Consult the paragraphs describing these instruments, and note particularly the difference in their construction.

Ex. 133. Because in the one saturated with moisture perspiration from the skin is impeded.

Ex. 134, pars. 215–218.

Ex. 135. Some of the air will rush out of the vessel on the top of the mountain; at the foot some of the external air will rush into the vessel.

Ex. 136. The surface of the earth and the air over it is much warmer than the surface of the sea and the air over it, during the day; the hot air over the island consequently ascends, and the cooler air, the *sea*-breeze, sets in from the sea. Conversely, during the night the sea gives out some of the heat it absorbed during the day, whilst the surface of the earth, from the absence of the sun, becomes comparatively cold. This reverse change in the temperature of the air over the land and sea produces the *land*-breeze.

Ex. 137, pars. 235–238.

Ex. 138. Because the air becomes more compressed, and therefore its sensible heat is diffused through less bulk.

Ex. 139, pars. 94, 104.

Ex. 140. The bubbles of vapour first formed *condense* on passing up through the colder layers of water; and the formation and condensation of the bubble occasion the *singing*.

Ex. 141. They are due to a current of moist air of a higher temperature than the surface of the sea; the cooling causes a condensation of some of the moisture, which appears in the form of fog or thick mist.

Ex. 142. Because they absorb moisture from the air; sulphuric acid is a liquid which absorbs a large quantity.

Ex. 143, par. 234.

Ex. 144. It is that temperature at which the tension of its vapour exactly balances the pressure of the atmosphere.

Ex. 145, par. 111.

Ex. 146. Because the motion of the body is not added to the muscular effort, as it is when a run is made before leaping.

Ex. 147, par. 237.

Ex. 148. On cloudy nights little or no dew is deposited, because the heat radiated from the earth is radiated back by the clouds, and therefore the reduction of temperature necessary for the deposition of dew does not take place; but on cloudless nights the heat emitted is not returned, hence dew is deposited.

Ex. 149, pars. 260, 261.

CHAPTER V.

Ex. 150. 0·86′.	Ex. 151. 0·8′.	Ex. 152. 1·175.
Ex. 153. 0·857.	Ex. 154. 9·53 lbs.	Ex. 155. 7·92 lbs.
Ex. 156. 11·72 lbs.	Ex. 157. 10·63 lbs.	
Ex. 158. 3430·9 grains.	Ex. 159. 2·5.	Ex. 160. 2·6′.
Ex. 161. 2·0.	Ex. 162. 5·8.	Ex. 163. 1·823.
Ex. 164. 3·6.	Ex. 165. 4·0.	Ex. 166. 2·143
Ex. 167. 4·824.	Ex. 168. 10·0.	Ex. 169. 0·844.
Ex. 170. 0·6.	Ex. 171. 0·777.	Ex. 172. 0·6.
Ex. 173. 1980 lbs.	Ex. 174. 1·45 cub. ft.	
Ex. 175. 486·3 lbs.	Ex. 176. 208·1 lbs.	Ex. 177. 52·8 lbs.

Ex. 178. 15 lbs. Ex. 179. 58·2 lbs. Ex. 180. 2644·2 grains.
Ex. 181. 437·5 lbs. Ex. 182. 3·73′ cub. feet.
Ex. 183. 1·5. Ex. 184. 0·79. Ex. 185. ·084.
Ex. 186. 1·16′ sp. gr. of the body; 0·875 sp. gr. of the spirit.
Ex. 187. 0·789. Ex. 188. 2·388. Ex. 189. 13·94.
Ex. 190. The alloy consists of 7·29 parts of gold, and 2·71 of silver.
Ex. 191. The alloy consists of 29·19 parts of silver, and 7·81 parts of copper.
Ex. 192. 17·608. Ex. 193. 16·04. Ex. 194. 1·555.
Ex. 199. 0·3420. Ex. 200. 0·3004. Ex. 201. 0·4729.
Ex. 202. 0·1825. Ex. 203. 71·54 cub. inches.
Ex. 204. 150·43 cub. inches. Ex. 205. 12·776 gram.
Ex. 206. 0·886 gram. Ex. 207. 19·54 gram.
Ex. 208. 3·25. Ex. 209. 1·547. Ex. 210. 1·07.
Ex. 211, pars. 292-294. Ex. 212, pars. 307, 308.
Ex. 213, par. 292. Ex. 214, par. 313.
Ex. 215. (1) par. 304; (2) par. 312; (3) par. 306.
Ex. 216, par. 320. Ex. 217. 4·22. Ex. 218. 5·83.
Ex. 219. No; cork, being of greater bulk, is buoyed up in the air by greater force, and hence weighs less in air; par. 324.
Ex. 220. 32 ft. 9·5 inches, par. 176. Ex. 221, par. 294.
Ex. 222. 0·8333.
Ex. 223. 105 pts. by wt. of silver to 29 pts. by wt. of copper.
Ex. 224, par. 180; oil barometer 11·692 metres.
Ex. 225, par. 309; 22·4 grms. in alcohol. Ex. 226, par. 250.
Ex. 227, par. 308; 0·84 sp. gr. of the spirit. Ex. 228, par. 108.
Ex. 229. 38 ft. 10·73 inches.
Ex. 330. Because it is buoyed up with greater force by the water.
Ex. 231. The weight of water displaced, and the quantity displaced, is equal in weight to the ship itself.
Ex. 232. As it is soluble in water, its sp. gr. must be determined in some liquid in which it is insoluble.
Ex. 233. Weight of a cub. ft. 450 lbs.; weight immersed in water 387·5 lbs.

CHAPTER VI.

Ex. 252. The chalk is insoluble, and the salt is soluble in water.

Ex. 253. The turbidity is due to the substance being less soluble in hot than in cold water.

Ex. 254. By evaporating the solution in an evaporating dish to dryness, and then examining whether any solid remains.

Ex. 255. When it ceases to dissolve it is a proof that the liquid is saturated, and by the solid dissolving in the boiling liquid it shows that it is more soluble in the hot than in the cold liquid.

CHAPTER VII.

Ex. 272, par. 381.

Ex. 273. By diminishing the atmospheric pressure the normal boiling-point of the ether is lowered, and it abstracts from the water the heat necessary for its ebullition, which causes the latter liquid to freeze.

Ex. 274. It is due to the gas expanding on removing the pressure.

Ex. 275. The ratio is as 0·27 to 1.

CHAPTER VIII.

Ex. 326, par. 410.

Ex. 327. The burning is due to their combining with some of the oxygen of the air.

Ex. 328, pars. 411, 412, 415.

Ex. 329. Expts. 280, 283–285; and 281, 282, 286.

Ex. 330, par. 432. Ex. 331, par. 420. Ex. 332, par. 438.

Ex. 333, pars. 407–409. Ex. 334, Expts. 286, 283.

Ex. 335, pars. 429, 430.

Ex. 336. Steam is water in the gaseous state, and its two elements, H and O, have entered into the highest state of combination.

Ex. 337, par. 445. Ex. 338, par. 449. Ex. 339, par. 447.

Ex. 340, par. 446. Ex. 341, par. 448. Ex. 342, par. 446.

Ex. 343, par. 440.

Ex. 344. No; it must be mixed with air or oxygen. The products of the explosion are carbon dioxide and water, and if the air has been employed there is likewise nitrogen.

Ex. 345, par. 459.

Ex. 346. Phosphoric anhydride (P_2O_5) and water (H_2O).

Ex. 347, par. 439.

Ex. 348. When burnt as it issues from the burner its elements only enter into union with the oxygen of the air at the point of contact, and therefore only in small quantities at a time; but when the air and gas are mixed the combination is instantaneous throughout the entire mass; and the great heat developed instantaneously causes a great and sudden expansion of the gaseous products. To this and the after-contraction the explosion is due.

Ex. 349. Expt. 292.

Ex. 350. According to the quantity of air, either the whole or part of the carbon will be set free in the form of soot.

Ex. 351, Expt. 322. Ex. 352, par. 434.

Ex. 353, pars. 345–347. Ex. 354, pars. 391, 407, 411, 459.

CHAPTER IX.

Ex. 366, par. 431. Ex. 367. It would simply volatilise.

Ex. 368. A gain in weight.

Ex. 369. Yes; the intensity would be increased.

Ex. 370. On account of the aqueous vapour it contains.

Ex. 371, par. 410. Ex. 372, par. 463.

Ex. 373, pars. 468, 469.

Ex. 375. Because the carbon cannot combine with a further quantity of oxygen.

Ex. 376, pars. 468, 469. Ex. 377, par. 462.

Ex. 378, par. 470.

Ex. 379. Because the hydrogen cannot combine with a further quantity of oxygen.

Ex. 380. Heavier by the amount of oxygen they have combined with.

Ex. 381, *a* 12·36 grams. 8·55 litres; *b* 18·39 grams. 12·75 litres.[1] Silver is extracted from argentiferous galena by the process described in par. 499.

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[1] The number of the Exercises ought to be *a*. 703; *b*. 664.

Ex. 382. The amount of iron oxidised would be 233·33 grams.; and the increase in weight would be 322·21 grms., and the volume of hydrogen gas set free would be 124 litres; for the formula of the oxide formed see par. 479.

Ex. 383. If we take ozone as the type of allotropic bodies in general, then we might infer that allotropy is due to the molecules of the different modifications of an element containing a different number of atoms.

Ex. 384. If the substance was exposed to heat and it took fire and burnt, it would prove it was diamond and not glass.

Ex. 385. It is most conveniently prepared by acting on MnO_2 with HCl; this is the way it is prepared on the manufacturing scale: 1225·3 grams. of MnO_2, and 2056·3 grams. of HCl.

Ex. 386. The earthy substances in ores are termed the *gangue*. *Slag* is the fusible mass that is formed by the fusion of the gangue with the substances added, which are called the *flux*; par. 495.

Ex. 387. 38·44 vols. of air; 2·0995 vols. of steam, 63·9 vols. of nitrogen in the air employed.

Ex. 388, par. 490. Ex. 389, par. 479. Ex. 390, par. 482.

Ex. 391. A reducing agent is a substance which removes oxygen from other substances, metals for instance, with which it is united, owing to its greater affinity for the oxygen. The reduction of metallic oxides on the manufacturing scale is also termed *smelting*.

Ex. 392, par. 487. Ex. 393, par. 503.

Ex. 394, pars. 491, 494, 505.

CHAPTER X.

Ex. 395. 31·5 lbs. of H_2O, and 1·5 lbs. of H in excess.

Ex. 396. 182·5 lbs. of HCl, and 0·5 lb. of Cl in excess.

Ex. 397. 896 lbs. of HI, and 111 lbs. of I in excess.

Ex. 398. 85 lbs. of H_2S, and 1 lb of H in excess.

Ex. 404. On account of not knowing the condensation that has taken place in the union of the O and N, we have therefore to ascertain how much H remains uncombined after the first explosion. We ascertain this by adding O and again exploding. There is a reduction in volume on this second explosion of 75 c.c., owing to the condensation of the water formed; therefore 50 c.c. of H remained uncombined after the first explosion; consequently 100 c.c. of the oxide of nitrogen gas consisted of 100 c.c. of N and 50 c.c. of O.

Ex. 405. 68·41 grains. Ex. 406. 327·65 cub. inches.

Ex. 407. 31·44 litres. Ex. 408. 0·716 gram.

Ex. 409. 69·92 litres; and 87·84 grams.

Ex. 410. 40 c.c. of aqueous vapour, and 10 c.c. of H.

Ex. 411. 500·07149 grams.

Ex. 412. One cubic inch; two cubic inches.

Ex. 413. One c.c.; two c.c. Ex. 414. 0·5 c.c.; 2 c.c.

Ex. 415. 0·9688. Ex. 416. 126·53. Ex. 417. One.

Ex. 418. 0·588. Ex. 419. 169·87. Ex. 420. Four.

Ex. 421. 1,176. Ex. 422. Two. Ex. 423. 9·39.

Ex. 424. 23·01 of O; and 76·99 of N. Ex. 425. 5,301.

Ex. 426. 106·5; three; $AuCl_3$. Ex. 427. 127; AuI.

Ex. 428. 24; Al_2O_3. Ex. 429. 106·5; BCl_3.

Ex. 430. 142·0; $SnCl_4$. Ex. 432. Na = 23, Ag = 108, Mg = 12.

Ex. 433. Ca = 20, B = 3·66, Sn = 29·5.

Ex. 434. $Fe'' = 28$, $Fe''' = 18{\cdot}66$.
Ex. 435, pars. 392, 576, 577, 520.   Ex. 436, par. 391.
Ex. 437, par. 541.   Ex. 438, par. 557.

CHAPTER XI.

Ex. 445, par. 584.   Ex. 446, par. 598.   Ex. 447, par. 593.
Ex. 465, par. 616—1st.
Ex. 466. Taking the limestone as pure, it would yield 1402·7 lbs.
Ex. 467. 154·9 lbs. of CaH_2O_2; and 35·6 lbs. of NH_3.
Ex. 468. 13 lbs.

CHAPTER XII.

Ex. 476. On account of the smaller atomic weight of sodium.
Ex. 477, par. 658.   Ex. 478, par. 643.
Ex. 479. ·053 grams of C.; 100 c.c. of CO_2.

Ex. 480. The acid acts upon the silver, converting it into sulphate, but it does not act upon the gold, par. 649. It is employed in separating gold and silver on the manufacturing scale. Concentrated acid has to be employed, and the alloy must not contain more than 25 per cent. of gold; if it contains more it must be further alloyed with a sufficient amount of silver.

Ex. 481, par. 573.   Ex. 482, par. 640.   Ex. 483, par. 643.
Ex. 484. 100 c.c. of each of the products.   Ex. 485, par. 649.
Ex. 491, par. 668—3rd.   Ex. 493, par. 642.   Ex. 494, par. 668—3rd.

Ex. 495. Phosphoric acid, because it is less volatile than sulphuric acid.

CHAPTER XIII.

Ex. 620. The hardness of waters is due to the presence of salts of lime and magnesia; their carbonates are insoluble in water, but they are dissolved in water containing CO_2, and on adding lime to water of this kind the carbonates are precipitated; see par. 742.

Ex. 621, par. 733.   Ex. 622, 36·5.   Ex. 623, pars. 720-722.
Ex. 624, pars. 735, 739.   Ex. 625. 0·16.
Ex. 626, par. 744.   Ex. 627, pars. 740, 738.
Ex. 628. 49.   Ex. 629, par. 746.

Ex. 630. A silicate will be formed, par. 668—3rd; on adding HCl the silicate will be decomposed, a chloride being formed, and a varying quantity of silicic acid will be precipitated, par. 668—2nd.

Ex. 631. 5·6.   Ex. 632. 0·84 gram.   Ex. 633. 32·66.

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Spottiswoode & Co. Printers, New-street Square, London.

NOV 3 - 1930